CHEMISTRY 142: PRE-GENERAL CHEMISTRY

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Chemistry 142: Pre-General Chemistry

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SECTION OVERVIEW

1.1: Introduction

1.1.1: The Scope of Chemistry

Learning Objectives

- To recognize the breadth, depth, and scope of chemistry.
- Define chemistry in relation to other sciences.
- Identify the main disciplines of chemistry.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—the screen on which you are reading this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter. Because of this, chemistry is a science that has its fingers in just about everything. Being able to describe the ingredients in a cake and how they change when the cake is baked, for example, is chemistry!

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 1.1.1.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

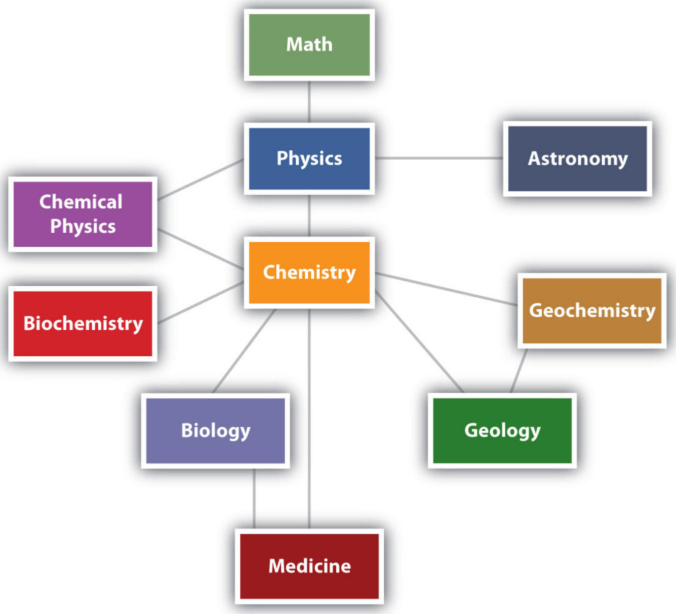

Figure 1.1.1.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 1.1.1.1 shows how many of the individual fields of science are related. At some level, all of these fields depend on matter because they all involve "stuff"; because of this, chemistry has been called the "central science", linking them all together.

> *There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed here.*

Example 1.1.1.1: Science Fields

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Solution

a. Sculpture is not considered a science because it is not a study of some aspect of the natural universe.
b. Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

? Exercise 1.1.1.1

Which fields of study are branches of science?

a. physiology (the study of the function of an animal's or a plant's body)
b. geophysics
c. agriculture
d. politics

Answer a:

yes

Answer b:

yes

Answer c:

yes

Answer d:

no

Areas of Chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study:

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- **Analytical chemistry:** Analytical chemistry is the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so that better treatments can be developed.

Figure 1.1.1.2: (left) Measurement of trace metals using atomic spectroscopy. (right) Measurement of hormone concentrations.

In practice, chemical research is often not limited to just one of the five major disciplines. A particular chemist may use biochemistry to isolate a particular chemical found in the human body such as hemoglobin, the oxygen carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to the areas of physical or analytical chemistry. Many chemists specialize in areas that are combinations of the main disciplines, such as bioinorganic chemistry or physical organic chemistry.

History of Chemistry

The **history** of chemistry is an interesting and challenging one. Very early chemists were often motivated mainly by the achievement of a specific goal or product. Making perfume or soaps did not need a lot of **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that we could all agree on). It is often difficult to figure out exactly what a particular person was using. However, the science developed over the centuries by trial and error.

Major progress was made toward putting chemistry on a solid foundation when Robert Boyle (1637-1691) began his **research** in chemistry (Figure $1.1.1.3$). He developed the basic ideas about the behavior of gases. He could then describe gases mathematically. Boyle also helped form the idea that small particles could combine to form molecules. Many years later, John Dalton used these ideas to develop the atomic theory.

Figure $1.1.1.3$: Robert Boyle.

The field of chemistry began to develop rapidly in the 1700's. Joseph Priestley (1733-1804) isolated and characterized several gases: oxygen, carbon monoxide, and nitrous oxide. It was later discovered that nitrous oxide ("laughing gas") worked as an anesthetic. This gas was used for that purpose for the first time in 1844 during a tooth extraction. Other gases discovered during that time were chlorine, by C.W. Scheele (1742-1786) and nitrogen, by Antoine Lavoisier (1743-1794). Lavoisier has been considered by many scholars to be the "father of chemistry". Among other accomplishments, he discovered the role of oxygen in combustion and definitively formulated the law of conservation of matter.

Chemists continued to discover new compounds in the 1800's. The science also began to develop a more theoretical foundation. John Dalton (1766-1844) put forth his atomic theory in 1807. This idea allowed scientists to think about chemistry in a much more systematic way. Amadeo Avogadro (1776-1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas. A lot of effort was put forth in studying chemical reactions. These efforts led to new materials being produced. Following the invention of the battery by Alessandro Volta (1745-1827), the field of electrochemistry (both theoretical and applications) developed through major contributions by Humphry Davy (1778-1829) and Michael Faraday (1791-1867). Other areas of the discipline also progressed rapidly.

It would take a large book to cover developments in chemistry during the twentieth century and up to today. One major area of expansion was in the area of the chemistry of living processes. Research in photosynthesis in plants, the discovery and characterization of enzymes as biochemical catalysts, elucidation of the structures of biomolecules such as insulin and DNA—these efforts gave rise to an explosion of information in the field of biochemistry.

The practical aspects of chemistry were not ignored. The work of Volta, Davy, and Faraday eventually led to the development of batteries that provided a source of electricity to power a number of devices (Figure $1.1.1.4$).

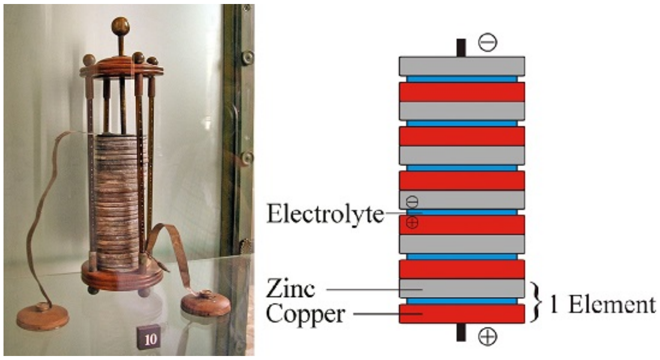

Figure 1.1.1.4: Battery developed by Volta. (CC BY-SA 3.0; (left) GuidoB and (right) Kkkdc).

Charles Goodyear (1800-1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles we have today. Louis Pasteur (1822-1895) pioneered the use of heat sterilization to eliminate unwanted microorganisms in wine and milk. Alfred Nobel (1833-1896) invented dynamite (Figure 1.1.1.5). After his death, the fortune he made from this product was used to fund the Nobel Prizes in science and the humanities. J.W. Hyatt (1837-1920) developed the first plastic. Leo Baekeland (1863-1944) developed the first synthetic resin, widely used for inexpensive and sturdy dinnerware.

Figure 1.1.1.5: Dynamite explosion in Panama, Central America (1908).

Today, chemistry continues to be essential to the development of new materials and technologies, from semiconductors for electronics to powerful new medicines, and beyond.

Summary

- Chemistry is the study of matter and the changes it undergoes and considers both *macroscopic* and *microscopic* information.
- Matter is anything that has mass and occupies space.
- The five main disciplines of chemistry are physical chemistry, organic chemistry, inorganic chemistry, analytical chemistry and biochemistry.
- Many civilizations contributed to the growth of chemistry. A lot of early chemical research focused on practical uses. Basic chemistry theories were developed during the nineteenth century. New materials and batteries are a few of the products of modern chemistry.

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1.1.2: Hypothesis, Theories, and Laws

Learning Objectives

- Describe the difference between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.

Although many have taken science classes throughout the course of their studies, people often have incorrect or misleading ideas about some of the most important and basic principles in science. Most students have heard of hypotheses, theories, and laws, but what do these terms really mean? Prior to reading this section, consider what you have learned about these terms before. What do these terms mean to you? What do you read that contradicts or supports what you thought?

What is a Fact?

A fact is a basic statement established by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

One of the most common terms used in science classes is a "hypothesis". The word can have many different definitions, depending on the context in which it is being used:

- An educated guess: a scientific hypothesis provides a suggested solution based on evidence.
- Prediction: if you have ever carried out a science experiment, you probably made this type of hypothesis when you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for it to be scientific, however, a scientist must be able to test the explanation to see if it works and if it is able to correctly predict what will happen in a situation. For example, "if my hypothesis is correct, we should see ___ result when we perform ___ test."

> *A hypothesis is very tentative; it can be easily changed.*

What is a Theory?

The *United States National Academy of Sciences* describes what a theory is as follows:

> "Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."
>
> "A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact.

Note some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomena. They aren't predictions (although we may use theories to make predictions). They are explanations as to why we observe something.
- Theories aren't likely to change. They have a large amount of support and are able to satisfactorily explain numerous observations. Theories can, indeed, be facts. Theories can change, but it is a long and difficult process. In order for a theory to change, there must be many observations or pieces of evidence that the theory cannot explain.
- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight; it is not just one person's idea about something

> *Theories aren't likely to change.*

What is a Law?

Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are more overarching explanations of how nature works and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas is increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

What is a Belief?

A **belief** is a statement that is not scientifically provable. Beliefs may or may not be incorrect; they just are outside the realm of science to explore.

Laws vs. Theories

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, **theories are explanations** and **laws are patterns** we see in large amounts of data, frequently written as an equation. A theory will always remain a theory; a law will always remain a law.

Video **1.1.2.1**: What's the difference between a scientific law and theory?

Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.

Contributions & Attributions

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- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

- **1.6: Hypothesis, Theories, and Laws** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

1.1.2.1: The Scientific Method - How Chemists Think

Learning Objectives

- Identify the components of the scientific method.

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments; which leads to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.1.2.1.1).

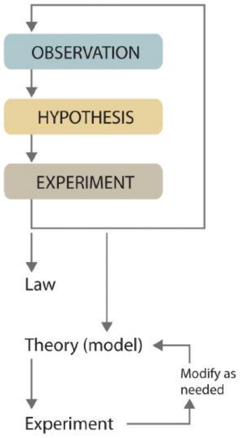

Figure 1.1.2.1.1 : The Steps in the Scientific Method.

Step 1: Make observations

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: "the outside air temperature is cooler during the winter season," "table salt is a crystalline solid," "sulfur crystals are yellow," and "dissolving a penny in dilute nitric acid forms a blue solution and a brown gas." **Quantitative observations** are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: "the melting point of crystalline sulfur is 115.21° Celsius," and "35.9 grams of table salt—the chemical name of which is sodium chloride—dissolve in 100 grams of water at 20° Celsius." For the question of the dinosaurs' extinction, the initial observation was quantitative: iridium concentrations in sediments dating to 66 million years ago were 20–160 times higher than normal.

Step 2: Formulate a hypothesis

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a hypothesis, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience alternating periods of light and darkness corresponding to observed movements of the sun, moon, clouds, and shadows is consistent with either one of two hypotheses:

a. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
b. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either support or refute it.

Step 3: Design and perform experiments

After a hypothesis has been formed, scientists conduct experiments to test its validity. Experiments are systematic observations or measurements, preferably made under controlled conditions—that is—under conditions in which a single variable changes.

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether or not the original hypothesis is valid. If the hypothesis is valid, the scientist can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified and requires further experimentation.

Step 5: Development into a law and/or theory

More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a law, a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply states what happens; it does not address the question of why.

One example of a law, the law of definite proportions, which was discovered by the French scientist Joseph Proust (1754–1826), states that a chemical substance always contains the same proportions of elements by mass. Thus, sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine, in this case 39.34% sodium and 60.66% chlorine by mass, and sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass.

Whereas a law states only what happens, a theory attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time unless a major experimental error is discovered. In contrast, a theory, by definition, is incomplete and imperfect, evolving with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure $1.1.2.1.1$ at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

✓ Example $1.1.2.1.1$

Classify each statement as a law, a theory, an experiment, a hypothesis, an observation.

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
d. When 10 g of ice were added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
e. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is a possible explanation for the origin of birds, so it is a hypothesis.
c. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
d. The temperature is measured before and after a change is made in a system, so these are observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer’s claim of purity), so it is an experiment.

? Exercise $1.1.2.1.1$

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really “consumes 47 times its weight in excess stomach acid.”
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter to ever play professional basketball.
e. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

observation

Summary

The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. The key steps in the scientific method include the following:

- Step 1: Make observations.
- Step 2: Formulate a hypothesis.
- Step 3: Test the hypothesis through experimentation.
- Step 4: Accept or modify the hypothesis.
- Step 5: Develop into a law and/or a theory.

Contributions & Attributions

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- Wikipedia

1.1.3: Extensive and Intensive Properties

Note from Dr. B.

Do not worry about the "Explore More" section. The link does not work.

How much is twenty dollars really worth?

You agree to mow someone's lawn for twenty dollars (it's a fairly large yard). When they pay you, they give you a $20 bill. It doesn't matter whether the bill is brand new or old, dirty, and wrinkled—all of these bills have the same value of $20. If you want more $20 bills, you have to mow more lawns. You can't say, "this particular bill is actually worth more than $20." To have more money, you have to put in more work.

Extensive Properties

Some properties of matter depend on the size of the sample, while some do not. An **extensive** property is a property that depends on the amount of matter in a sample. The mass of an object is a measure of the amount of matter that an object contains. A small sample of a certain type of matter will have a small mass, while a larger sample will have a greater mass. Another extensive property is volume. The volume of an object is a measure of the space that is occupied by that object.

The figure below illustrates the extensive property of volume. The pitcher and glass both contain milk. The pitcher holds approximately two quarts and the glass will hold about 8 ounces of milk. The same milk is in each container. The only difference is the amount of milk contained in the glass and in the pitcher.

Figure 1.1.3.1: Milk pitcher and glass. (Credit: Zenon Niewada (Wikimedia: Pitcherman); Source: http://commons.wikimedia.org/wiki/File:Milk_Pitcher_With_Lid.jpg(opens in new window); License: Public Domain)

Intensive Properties

The electrical conductivity of a substance is a property that depends only on the type of substance. Silver, gold, and copper are excellent conductors of electricity, while glass and plastic are poor conductors. A larger or smaller piece of glass will not change this property. An **intensive** property is a property of matter that depends only on the type of matter in a sample and not on the amount. Other intensive properties include color, temperature, density, and solubility.

Figure 1.1.3.1: *Copper wire.* (Credit: User:Inductiveload/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Tinned_Copper_Wire_anaglyph.jpg(opens in new window))

The copper wire shown in the picture below has a certain electrical conductivity. You could cut off the small end that sticks out, and it would have the same conductivity as the entire long roll of wire shown here. The conductivity is a property of the copper metal itself, not of the length of the wire.

Summary

- An extensive property is a property that depends on the amount of matter in a sample.
- Mass and volume are examples of extensive properties.
- An intensive property is a property of matter that depends only on the type of matter in a sample and not on the amount.
- Color, temperature, and solubility are examples of intensive properties.

Review

1. Define extensive property.
2. Give two examples of extensive properties.
3. Define intensive property.
4. Give two examples of intensive properties.

1.1.4: Measurements and Exact Numbers

Measurement

Measurements have three parts to them:

1. Numeric Value
2. Unit
3. Uncertainty

For example, if a certain length is measured as 45.2 cm, the numerical value is 45.2 and the unit is cm (centimeter). The last significant digit reported communicates the level of uncertainty. Since the last digit in this example is 2, that is the uncertain digit. The 2 is uncertain and could possibly be 1 or 3. So this length is somewhere between 45.1 cm and 45.3 cm.

If instead the length had been reported as 45 cm, the 5 would be the uncertain digit, and the measurement is between 44 cm and 46 cm.

Both of these measurements might have been taken of the same length, but one of them was taken with a more precise instrument and one was taken with a less precise instrument. Reporting one of these measurements tells us not only the length itself, but how certain we are of the measurement and the precision of the instrument used to take the measurements.

Note

Any measured value has a component of uncertainty, no matter how precise the instrument used. The convention used is that the last significant digit reported should represent the place value of the uncertainty. That is, the last (right-most) digit is the uncertain one!

Exact Numbers

Exact numbers have no uncertainty. You can think of them as having an infinite number of significant digit. This is not possible for any measurement, but it is possible in these cases:

1. Counted items
2. Defined values

For example, if you pour a substance into four beakers, the number 4 is exact. There is no uncertainty in the number of beakers you have. Also, if you measure the diameter of a circle, and you calculate the radius by dividing the diameter by 2, the number 2 is exact. There are exactly two radii in the diameter.

Also, some values are exact because one unit is defined as being exactly equal to another. Any metric relationship is exact. There are exactly 1000 m in 1 km. For some relationships, it is not as obvious that the relationship is exact, and you will be informed in the textbook or in the problem. For example, the inch is defined as being equal to 2.54 cm exactly.

[Note that not all relationships we will use are exact. Some are approximations that have some uncertainty. One example is the use of the relationship "1 kg = 2.20 lb". The kilogram is not defined as 2.20 pounds. It is not exactly 2.20 pounds. This is an approximation. If we want to use a more precise approximation, we would use "1 kg = 2.2046 lb". If we wanted to be even more precise, we could use "1 kg = 2.2046226 lb". We could even look in a reference book to find more digits than that, but these approximations are good enough for our purposes.]

Exercise 1.1.4.1

Determine whether each number is exact or measured:

(a) The homework assignment has 8 questions.

(b) The assignment took 15 minutes to complete.

(c) There are 60 minutes in an hour.

(d) Therefore the homework assignment took 0.25 hours.

Answer (a)

Exact.

Answer (b)

Measured.

Answer (c)

Exact.

Answer (d)

Measured.

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1.1.5: Accuracy and Precision

Figure 1.1.5.1 (Credit: Daniel Arizpe; Source: Commons Wikimedia, High School Basketball Game(opens in new window) [commons.wikimedia.org]; License: Public Domain)

How do professional basketball players improve their shooting accuracy?

Basketball is one of those sports where you need to hit the target. A football field goal kicker might have room for some deviation from a straight line – for college and pro football there is an 18 foot 6 inch space for the ball to go through. In basketball, the basket is only 18 inches across and the ball is a little less than 10 inches across – not much room for error. The ball has to be on target in order to go into the basket and score.

Accuracy and Precision

In everyday speech, the terms **accuracy** and **precision** are frequently used interchangeably. However, their scientific meanings are quite different. Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured. Precision is a measure of how close a series of measurements are to one another. Precise measurements are highly reproducible, even if the measurements are not near the correct value. Darts thrown at a dartboard are helpful in illustrating accuracy and precision.

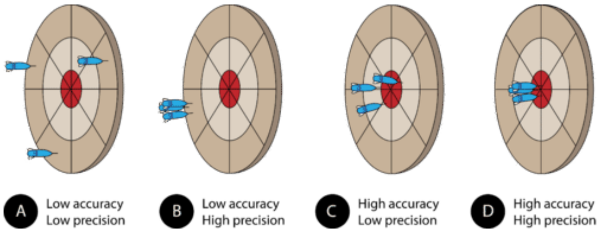


Figure 1.1.5.2: The distribution of darts on a dartboard shows the difference between accuracy and precision. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. The figure above demonstrates four possible outcomes:

a. The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate nor precise.
b. The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly, or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.
c. The darts are not grouped very near to each other, but are generally centered around the bulls-eye. This demonstrates poor precision, but fairly high accuracy. This situation is not desirable in a lab situation because the "high" accuracy may simply be random chance and not a true indicator of good measuring skill.
d. The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements.

Figure 1.1.5.3: Students in a chemistry lab are making careful measurements with a series of volumetric flasks. Accuracy and precision are critical in every experiment. (Credit: CK-12 Foundation; Source: CK-12 Foundation; License: CK-12 Curriculum Materials license)

Summary

- Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured.
- Precision is a measure of how close a series of measurements are to one another.

Review

1. Define accuracy.
2. Define precision.
3. What can be said about the reproducibility of precise values?

1.1.5.1: Percent Error

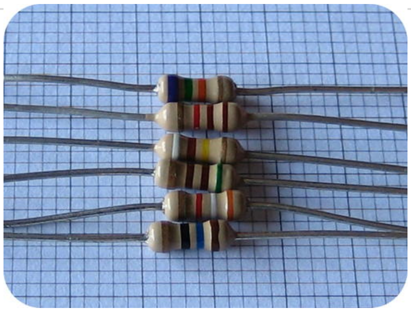


Figure 1.1.5.1.1 (Credit: Vald Kliper; Source: http://commons.wikimedia.org/wiki/File:Resistencias_250W_5%2525_sobre_papel_milimetrado.JPG(opens in new window); License: Public Domain)

How does an electrical circuit work?

A complicated piece of electronic equipment may contain several resistors that have the role of controlling the voltage and current in the electrical circuit. If there is too much current, the apparatus malfunctions. Conversely, too little current means the system simply will not perform. The resistor values are always given with an error range. A resistor may have a stated value of 200 ohms, but a 10% error range, meaning the resistance could be anywhere between 195 and 205 ohms. By knowing these values, an electronics person can design and service the equipment to make sure that it functions properly.

Percent Error

An individual measurement may be accurate or inaccurate, depending on how close it is to the true value. Suppose that you are doing an experiment to determine the density of a sample of aluminum metal. The **accepted value** of a measurement is the true or correct value based on general agreement with a reliable reference. For aluminum, the accepted density is 2.70 g/cm^3. The **experimental value** of a measurement is the value that is measured during the experiment. Suppose that in your experiment, you determine an experimental value for the aluminum density to be 2.42 g/cm^3. The **error** of an experiment is the difference between the experimental and accepted values.

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

If the experimental value is less than the accepted value, the error is negative. If the experimental value is larger than the accepted value, the error is positive. Often, error is reported as the absolute value of the difference in order to avoid the confusion of a negative error. The **percent error** is the absolute value of the error, divided by the accepted value, and multiplied by 100%.

$$\%\text{ Error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

To calculate the percent error for the aluminum density measurement, we can substitute the given values of 2.45 g/cm^3 for the experimental value and 2.70 g/cm^3 for the accepted value.

$$\%\text{ Error} = \frac{|2.45 \text{ g/cm}^3 - 2.70 \text{ g/cm}^3|}{2.70 \text{ g/cm}^3} \times 100\% = 9.26\%$$

If the experimental value is equal to the accepted value, the percent error is equal to 0. As the accuracy of a measurement decreases, the percent error of that measurement rises.

Summary

- Definitions of accepted value and experimental values are given.
- Calculations of error and percent error are demonstrated.

Review

1. Define accepted value.
2. Define experimental value
3. What happens as the accuracy of the measurement decreases?

1.1.5.2: Comprehension Check

Exercise 1.1.5.2.1

A reported value of 45.35 g implies that the mass is between ___ and ___.

Answer

Mass is between 45.34 g and 45.36 g. It is 45.35 g plus or minus 0.01 g.

Is it possible for a number to have NO uncertainty?

Answer

A defined value or a number of counted items could be exact and therefore have no uncertainty.

Which of the following could be used to measure the precision of a set of numbers?

a) average (mean)

b) percent error

c) standard deviation

d) median

Answer

c) standard deviation (It tells how close a group of values are to each other. The other choices do not do that.)

Exercise 1.1.5.2.1

You perform an experiment to determine the density of copper, and you obtain a value of 8.62 g/mL. Your reference lists the accepted value as 8.96 g/mL. What is the percent error of your measured value?

Answer

It is 3.8%.

The difference between the values, 0.34 g/mL, is divided by the accepted value of 8.96 g/mL. The result is 0.038. This is multiplied by 100% to give 3.8%.

Exercise 1.1.5.2.1

You are measuring the liquid that is dispensed by a pump that is supposed to dispense 25.00 mL. You measure the following amounts: 24.72 mL, 24.88 mL, 24.85 mL and 24.77 mL. Calculate the average, standard deviation, and percent error.

Answer

Average = 24.81 mL. Standard deviation = 0.07 mL. Percent error = 0.76 %

1.1.6: SI Base Units

How long is a yard?

It depends on whom you ask, and when you asked the question. Today we have a standard definition of the yard, which you can see marked on every football field. If a football player moves the ball ten yards, they get a first down, and it doesn't matter whether they are playing in Los Angeles, Dallas, or Green Bay. At one time, however, that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. Of course, the problem in that case is simple: new king, new distance (and then you have to re-mark all those football fields).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units, listed in Table 1.1.6.1.

Table 1.1.6.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Figure 1.1.6.1: Meter Standard. (Credit: Courtesy of US Department of Commerce; Courtesy of National Institute of Standards and Technology; Source: http://commons.wikimedia.org/wiki/File:US_National_Length_Meter.JPG(opens in new window); http://commons.wikimedia.org/wiki/File:Standard_kilogram,_2.jpg(opens in new window); License: Public Domain)

Figure 1.1.6.2: Kilogram standard (Credit: Courtesy of National Institute of Standards and Technology; Source: http://commons.wikimedia.org/wiki/File:Standard_kilogram,_2.jpg(opens in new window); License: Public Domain)

The map below shows the adoption of the SI units in countries around the world. The United States has legally adopted the metric system for measurements, but does not use it in everyday practice. Great Britain and much of Canada use a combination of metric and imperial units.

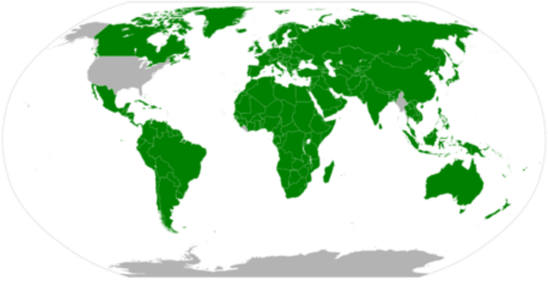


Figure 1.1.6.3: Areas of the world that use the metric system (green). (Credit: User:Canuckguy/Wikipedia and User:AzaToth/Wikimedia Commons;
Source: http://commons.wikimedia.org/wiki/File:Metric_system_adoption_map.svg(opens in new window); License: Public Domain)

Summary

- The SI system is based on multiples of ten.
- There are seven basic units in the SI system.
- Five of these seven SI units are commonly used in chemistry.

Review

1. What does SI stand for?
2. When was this system adopted by the international community?
3. Which of the units are commonly used in chemistry?

1.1.6.1: SI Length and Volume Units

Note from Dr. B.

Do not worry about the "Explore More" section. The link does not work.

Figure 1.1.6.1.1 (Credit: User:Żeglarz/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Prawa_burta_ORP_Iskra.JPG(opens in new window); License: Public Domain)

How were sailors able to measure the depths of seas?

Back in the days before all the electronic gadgets for measuring depth and locating undersea objects existed, the "fathom" was the unit of measurement for depth. A rope was knotted every six feet and the end was dropped over the side of the ship. You could tell how deep the water was by how many knots went under the water before the rope hit bottom. Today we just turn on an instrument and read the depth to a high level of accuracy.

Length and Volume

Length is the measurement of the extent of something along its greatest dimension. The SI basic unit of length, or linear measure, is the **meter** (m). All measurements of length may be made in meters, though the prefixes listed in various tables will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance, such as the distance between New York City and Chicago, is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 20 micrometers $(\mu\mathrm{m})$.

Volume is the amount of space occupied by a sample of matter. The volume of a regular object can be calculated by multiplying its length by its width and height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter $\left(\mathrm{m}^3\right)$, which is the volume occupied by a cube that measures $1 \mathrm{~m}$ on each side. This very large volume is not convenient for typical use in a chemistry laboratory. A liter (L) is the volume of a cube that measures $10 \mathrm{~cm}$ $(1 \mathrm{~dm})$ on each side. A liter is thus equal to both $1000 \mathrm{~cm}^3$ $(10 \mathrm{~cm} \times 10 \mathrm{~cm} \times 10 \mathrm{~cm})$, and to $1 \mathrm{~dm}^3$. A smaller unit of volume that is commonly used is the milliliter (mL—note the capital L, which is a standard practice). A milliliter is the volume of a cube that measures $1 \mathrm{~cm}$ on each side. Therefore, a milliliter is equal to a cubic centimeter $\left(\mathrm{cm}^3\right)$. There are $1000 \mathrm{~mL}$ in $1 \mathrm{~L}$, which is the same as saying that there are $1000 \mathrm{~cm}^3$ in $1 \mathrm{~dm}^3$.

Figure 1.1.6.1.2: A typical water bottle is 1 liter in volume. (Credit: User:Kenyon/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:CamelBak_water_bottle.jpg(opens in new window); License: Public Domain)

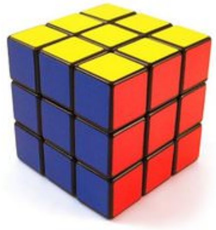


Figure 1.1.6.1.3: This Rubik's cube is 5.7 cm on each side and has a volume of 185.2 cm^3 or 185.2 mL. (Credit: Bram Van Damme (Flickr: Bramus!); Source: http://www.flickr.com/photos/bramus/3249196137/(opens in new window); License: CC by 2.0(opens in new window))

Figure 1.1.6.1.4: A graduated cylinder is often used to measure volume in the laboratory, and comes in a variety of sizes. (Credit: User:Darrien/Wikipedia; Source: http://commons.wikimedia.org/wiki/File:Graduated_cylinder.jpg(opens in new window); License: Public Domain)

Summary

- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of an item.

Review

1. Define length.
2. Define volume.
3. An object measures 6.2 cm×13.7 cm×26.9 cm6.2 cm×13.7 cm×26.9 cm. Which value is the length of the object?
4. How big is a mL?

1.1.6.2: SI Mass and Weight Units

Note from Dr. B.

Do not worry about the "Explore More" section. The link doesn't work.

Figure 1.1.6.2.1 (Credit: Courtesy of NASA; Source: http://commons.wikimedia.org/wiki/File:151746main_s121e05215-lg.jpeg(opens in new window); License: Public Domain)

How is he floating?

One of the many interesting things about travel in outer space is the idea of weightlessness. If something is not fastened down, it will float in mid-air. Early astronauts learned that weightlessness had negative effects on bone structure. If there was no pressure on the legs, those bones would begin to lose mass. Weight provided by gravity is needed to maintain healthy bones. Specially designed equipment is now a part of every space mission, so the astronauts can maintain good body fitness.

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The base SI unit of mass is the kilogram or kg, which was originally defined as the mass of 1 L of liquid water at 4^oC (the volume of a liquid changes slightly with temperature). The kilogram is the only SI base unit that contains a prefix. This is because when the first physical standard for mass was created in 1799, the methods to measure the mass of something as small as a gram did not exist. In the laboratory, mass is measured with a balance (figure below), which must be calibrated with a standard mass so that its measurements are accurate.

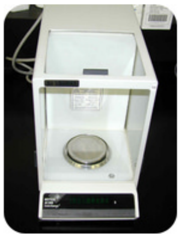


Figure 1.1.6.2.2: An analytical balance makes very sensitive mass measurements in a laboratory, usually in grams. (Credit: Courtesy of US Drug Enforcement Administration; Source: http://commons.wikimedia.org/wiki/File:Analytical_balance_mettler_ae-260.jpg(opens in new window); License: Public Domain)

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are 1000 g in 1 kg. A milligram is 1/1000th of a gram, so there are 1000 mg in 1 g.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to **gravity** is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale by reading the distance that a certain object pulls down and stretches a spring.

Summary

- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.

Review

1. Define mass.
2. Define weight.
3. If I weigh 180 pounds on Earth, what will I weigh on the moon?

1.1.6.3: Derived Units

Figure 1.1.6.3.1 (Credit: User:Joegrimes/Wikipedia; Source: Commons Wikimedia, Farm landly sander(opens in new window) [commons.wikimedia.org]; License: Public Domain)

How has farming evolved?

As farming becomes more expensive and less profitable (at least for small farms), farmers might decide to sell their land to builders who want to erect commercial or residential properties. In order to sell, an accurate property title is needed. The dimensions of the farm must be determined, and the acreage calculated from those dimensions.

Dimensional Analysis and Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. We have already discussed volume and energy as two examples of derived units. Some others are listed in the table below:

Derived SI Units

Quantity	**Symbol**	**Unit**	**Unit Abbreviation**	**Derivation**
Area	A	square meter	m^2	leng
Volume	V	cubic meter	m^3	leng

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Density	D	kilograms/cubic meter	kg/	`\(\frac{\text{mass}}{\text{volume}}\0`
Concentration	c	moles/liter	mol	amou volu
Speed (velocity)	v	meters/second	m/s	lengt time
Acceleration	a	meters/second/second	m/s	speed time
Force	F	newton	N	mas
Energy	E	joule	J	forc

Using dimensional analysis with derived units requires special care. When units are squared or cubed as with area or volume, the conversion factors themselves must also be squared. Shown below is the conversion factor for cubic centimeters and cubic meters.

$$\left(\frac{1\ \text{m}}{100\ \text{cm}}\right)^3 = \frac{1\ \text{m}^3}{10^6\ \text{cm}^3} = 1$$

Because a cube has 3 sides, each side is subject to the conversion of 1 m to 100 cm. Since 100 cubed is equal to 1 million $\left(10^6\right)$, there are 10^6 cm^3 in 1 m^3. Two convenient volume units are the liter (which is equal to a cubic decimeter) and the milliliter, which is equal to a cubic centimeter. The conversion factor would be:

$$\left(\frac{1 \text{ dm}}{10 \text{ cm}}\right)^3 = \frac{1 \text{ dm}^3}{1000 \text{ cm}^3} = 1$$

There are thus 1000 cm^3 in 1 dm^3, which is the same thing as saying there are 1000 mL in 1 L.

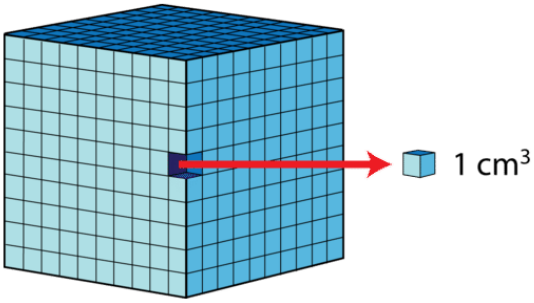


Figure 1.1.6.3.2: There are 1000 cm^3 in 1 dm^3. Since a cm^3 is equal to a mL and a dm^3 is equal to a L, we can say that there are 1000 mL in 1 L. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Example 1.1.6.3.1

Convert $3.6 \times 10^8 \text{ mm}^3$ to mL.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1 \text{ m} = 1000 \text{ mm}$
- $1 \text{ mL} = 1 \text{ cm}^3$
- $1 \text{ m} = 100 \text{ cm}$

Unknown

This problem requires multiple steps and the technique for converting with derived units. Simply proceed one step at a time: mm^3 to m^3 to $\text{cm}^3 = \text{mL}$.

Step 2: Calculate.

$$3.6 \text{ mm}^3 \times \left(\frac{1 \text{ m}}{1000 \text{ mm}}\right)^3 \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 0.0036 \text{ mL}$$

Numerically, the steps are to divide 3.6 by 10^9, followed by multiplying by 10^6. You may find that you can shorten the problem by a step by first determining the conversion factor from mm to cm, and using that instead of first converting to m. There are 10 mm in 1 cm.

$$3.6 \text{ mm}^3 \times \left(\frac{1 \text{ cm}}{10 \text{ mm}}\right)^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 0.0036 \text{ mL}$$

In this case, 3.6/1000 gives the same result of 0.0036.

Step 3: Think about your result.

Cubic millimeters are much smaller than cubic centimeters, so the final answer is much less than the original number of mm^3.

Summary

- A derived unit is a unit that results from a mathematical combination of SI base units.
- Calculations involving derived units follow the same principles as other unit conversion calculations.

Review

1. What is a derived unit?
2. Convert 0.00722 km^3 to m^3
3. Convert 129 cm^3 to L
4. Convert 4.9 × 105 μm^3 to mm^3

- **3.10: Derived Units** by CK-12 Foundation is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

1.1.6.4: SI Kinetic Energy Units

Figure 1.1.6.4.1 (Credit: User:Dlls/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Cat_public_domain_dedication_image_0002.jpg(opens in new window); License: Public Domain)

Have you ever watched a cat in action?

When a cat is chasing something, its moves very fast. A person may comment, "that cat has a lot of energy"—which is more correct than they might realize! One form of energy is seen when an object is moving, and this type of energy is the basis for many chemical processes.

SI Kinetic Energy Units

An object's **kinetic energy** is the energy due to motion. Kinetic energy can be defined mathematically as:

$$KE = \frac{1}{2}mv^2$$

where KE is kinetic energy, m is mass, and v is velocity.

Energy is defined as the capacity to do work or to produce heat. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered energy is **potential energy**, a type of energy that is stored in matter and released during a chemical reaction. The **joule** (J) is the SI unit of energy and is named after English physicist James Prescott Joule (1818-1889). If we go back to the equation for kinetic energy written above, we can put units in: kg for mass and $\mathrm{m^2/s^2}$ for velocity squared. Then, in terms of SI base units, a joule is equal to a kilogram times meter squared, divided by a second squared $(\mathrm{kg \cdot m^2/s^2})$. Another common unit of energy that is often used is the **calorie** (cal), which is equivalent to $4.184 \mathrm{~J}$.

Figure 1.1.6.4.2: James Prescott Joule. (Credit: John Collier; Source: http://commons.wikimedia.org/wiki/File:James_Prescott_Joule_by_John_Collier,_1882.jpg(opens in new window); License: Public Domain)

Summary

- Energy is the capacity to do work or to produce heat.
- Kinetic energy is energy due to motion.
- ~~Potential energy is energy stored in matter.~~
- The joule (J) is the SI unit of energy and equals $\mathrm{kg} \cdot \mathrm{m}^2/\mathrm{s}^2$.

Review

1. What is kinetic energy?
2. What is the mathematical equation for kinetic energy?
3. What is potential energy?
4. What is the SI unit for energy?

SECTION OVERVIEW

1.2: Numbers and Measurement

1.2.1: Significant Figures - Writing Numbers to Reflect Precision

Note from Dr. B.

Rule 4 for significant figures below says that trailing zeros are ambiguous and that scientific notation should be used. This is the best way to show what you mean. If there are trailing zeroes and no scientific notation, we will use the rule "depends on decimal". If there is no decimal shown, we must assume the zeros are not significant. So for the ambiguous examples below, we would say that 1400 has 2 significant figures, 500 has 1 significant figure, and 3800 has 2 significant figures.

Also, I do not completely agree with the archery target example below. See what you think, and use it as a topic for discussion this week!

Learning Objectives

- Identify the number of significant figures in a reported value.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length at 2.553 cm as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

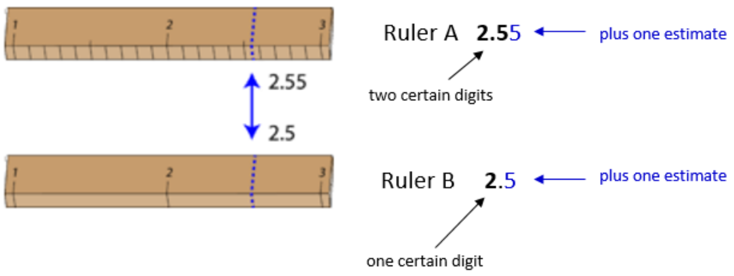


Figure 1.2.1.1: Measurement with two different rulers.
Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

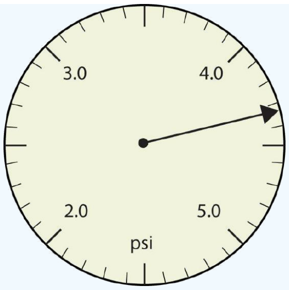
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure $1.2.1.1$ shows two rulers making the same measurement of an object (indicated by the blue arrow).

With either ruler, it is clear that the length of the object is between 2 and 3 cm. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as 2.5 cm. However, another person may judge that the measurement is 2.4 cm or perhaps 2.6 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

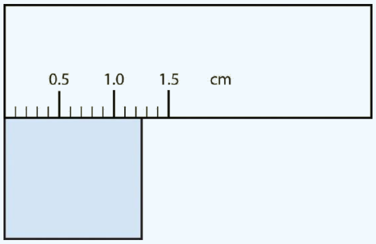
✓ Example 1.2.1.1: Reporting Measurements to the Proper Number of Significant Figures

Use each diagram to report a measurement to the proper number of significant figures.



a.

b.



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

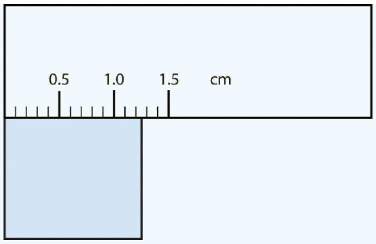
Solutions

Solutions to Example 2.3.1

	Explanation	Answer
a.	The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it's at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. The symbol psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire. The measurement is reported to three significant figures.	4.33 psi
b.	The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle's width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. The measurement is reported to three significant figures.	1.25 cm

Exercise 1.2.1.1

What would be the reported width of this rectangle?



Answer

1.25 cm

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 1.2.1.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous. ∘ 1.4×10^3 has two significant figures. ∘ 1.40×10^3 three significant figures. ∘ 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000…). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

✓ Example 1.2.1.2

Give the number of significant figures in each. Identify the rule for each.

a. 5.87
b. 0.031
c. 52.90
d. 00.2001
e. 500
f. 6 atoms

Solution

Solution to Example 2.3.2

Explanation	Answer

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d	The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e	The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f	The 6 is a counting number. A counting number is an exact number.	6, infinite

Exercise 1.2.1.2

Give the number of significant figures in each.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people
e. 3,800 g

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

Answer e

Ambiguous, could be two, three or four significant figures.

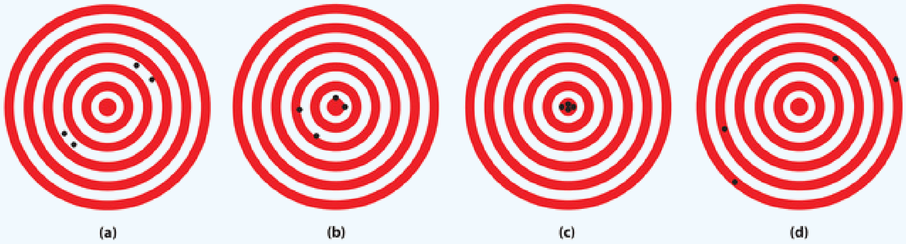
Accuracy and Precision

Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrates the concepts of accuracy and precision.

Video 1.2.1.1: Difference between precision and accuracy.

Example 1.2.1.3

The following archery targets show marks that represent the results of four sets of measurements.



(a) (b) (c) (d)

Which target shows

a. a precise, but inaccurate set of measurements?
b. a set of measurements that is both precise and accurate?
c. a set of measurements that is neither precise nor accurate?

Solution

a. Set **(a)** is precise, but inaccurate.
b. Set **(c)** is both precise and accurate.
c. Set **(d)** is neither precise nor accurate.

Summary

Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool. Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

1.2.2: Significant Figures in Calculations

Note from Dr. B.

It is not obvious in the last example and exercise that the number 2 and the number 5 are meant to be exact numbers. In problems that I write, I would specify that a number is counted or exact. If it is not specified, you should assume that the number 2 has only one significant figure.

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement $207.518\ \mathrm{m}$ Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table $1.2.2.1$.

Table $1.2.2.1$: Rounding examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Notice that the more rounding that is done, the less reliable the figure is. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

It is important to be aware of significant figures when you are mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238… to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers that have only three significant figures each? When performing mathematical operations, there are two rules for limiting the number of significant figures in an answer—one rule is for addition and subtraction, and one rule is for multiplication and division.

> *In operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the **least precise** operation. An answer is no more precise than the least precise number used to get the answer.*

Multiplication and Division

For multiplication or division, the rule is to count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

$$\underbrace{38.65}_{\text{4 sig figs}} \times \underbrace{105.93}_{\text{5 sig figs}} = \underbrace{4{,}094.1945}_{\text{reduce to 4 sig figs}}$$

The final answer, limited to four significant figures, is 4,094. The first digit dropped is 1, so we do not round up.

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 450 has two significant figures and would be written in scientific notation as 4.5×10^2, whereas 450.0 has four significant figures and would be written as 4.500×10^2. In scientific notation, all significant figures are listed explicitly.

✓ Example 1.2.2.1

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 23.096 × 90.300
b. 125 × 9.000

Solution

a

Explanation	Answer
The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	2.0856×10^3

b

Explanation	Answer
The calculator gives 1,125 as the answer, but we limit it to three significant figures.	1.13×10^3

Addition and Subtraction

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.71, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

$$\begin{array}{l} 1.2 \\ \underline{4.41} \\ 5.61 \\ \ \uparrow \text{ limit final answer to the tenths column: } 5.6 \end{array}$$

We drop the last digit—the 1—because it is not significant to the final answer.

The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater, and rounded down if the first dropped digit is less than 5.

$$\begin{array}{l} 77.2 \\ \underline{10.46} \\ 87.66 \\ \ \ \ \uparrow \text{ limit final answer to the tenths column and round up: } 87.7 \end{array}$$

✓ Example 1.2.2.2

a. 13.77 + 908.226
b. 1,027 + 611 + 363.06

Solution

a

Explanation	Answer
The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place, we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$

b

Explanation	Answer
The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2{,}001.06 = 2.001 \times 10^3$

? Exercise 1.2.2.2

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 217 ÷ 903
b. 13.77 + 908.226 + 515
c. 255.0 − 99
d. 0.00666 × 321

Answer a:

$$0.240 = 2.40 \times 10^{-1}$$

Answer b:

$$1,437 = 1.437 \times 10^3$$

Answer c:

$$156 = 1.56 \times 10^2$$

Answer d:

$$2.14 = 2.14 \times 10^0$$

Remember that calculators do not understand significant figures. *You* are the one who must apply the rules of significant figures to a result from your calculator.

Calculations Involving Multiplication/Division and Addition/Subtraction

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain at least one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.

Video 1.2.2.1: Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).

Video 1.2.2.2: https://www.youtube.com/watch?v=__csP0NtlGI

In the worked examples in this text, we will often show the results of intermediate steps in a calculation. In doing so, we will show the results to only the correct number of significant figures allowed for that step, in effect treating each step as a separate calculation. This procedure is intended to reinforce the rules for determining the number of significant figures, but in some cases it may give a final answer that differs in the last digit from that obtained using a calculator, where all digits are carried through to the last step.

✓ Example 1.2.2.3

a. 2(1.008 g) + 15.99 g
b. 137.3 s + 2(35.45 s)
c. $\frac{118.7g}{2}-35.5g$

Solution

a.

Explanation	Answer

2(1.008 g) + 15.99 g = **Perform multiplication first.** 2 (1.008 g *4 sig figs*) = **2.016 g** *4 sig figs* The number with the least number of significant figures is 1.008 g; *the number 2 is an exact number and therefore has an infinite number of significant figures*. **Then, perform the addition.** **2.016 g** *thousandths place* + 15.99 g *hundredths place (least precise)* = 18.006 g **Round the final answer.** Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)

b.

Explanation	Answer
137.3 s + 2(35.45 s) = **Perform multiplication first.** 2(35.45 s *4 sig figs*) = **70.90 s** *4 sig figs* The number with the least number of significant figures is 35.45; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Then, perform the addition.** 137.3 s *tenths place (least precise)* + **70.90 s** *hundredths place* = 208.20 s **Round the final answer.** Round the final answer to the tenths place based on 137.3 s.	208.2 s

c.

Explanation	**Answer**
$\frac{118.7g}{2} - 35.5g =$ **Perform division first.** $\frac{118.7g}{2}$ *4 sig figs* = **59.35 g** *4 sig figs* The number with the least number of significant figures is 118.7 g; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Perform subtraction next.** **59.35 g** *hundredths place* − 35.5 g *tenths place (least precise)* = 23.85 g **Round the final answer.** Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

? Exercise 1.2.2.3

Complete the calculations and report your answers using the correct number of significant figures.

a. 5(1.008s) - 10.66 s
b. 99.0 cm+ 2(5.56 cm)

Answer a

-5.62 s

Answer b

110.2 cm

Summary

- Rounding
    - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1 (e.g. 2.9699 rounded to three significant figures is 2.97).
    - If the number to be dropped is less than 5, there is no change (e.g. 4.00443 rounded to four significant figures is 4.004).
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the number with the fewest significant figures.
- The rule in addition and subtraction is that the answer is given the same number of decimal places as the term with the fewest decimal places.

1.2.2.1: Practice with Significant Figures

Exercise 1.2.2.1.1

Tell the number of significant figures in each measurement.

(a) 5060 L

(b) 340 g

(c) 3 cm

(d) 30.04 mg

(e) 0.00440 kg

(f) 640.0 mL

Answer

(a) 3

(b) 2

(c) 1

(d) 4

(e) 3

(f) 4

Exercise 1.2.2.1.1

Round each number to two significant digits.

(a) 4.825

(b) 0.0568

(c) 5624.8

Answer

(a) 4.8

(b) 0.057

(c) 5600

Round each number to the tens place.

(d) 118.8

(e) 49.3

(f) 3582

Answer

(d) 120

(e) 50

(f) 3580

Exercise 1.2.2.1.1

Provide the answer to each multiplication or division problem with correct significant digits.

(a) 3.4 X 27.8

(b) 6600/2.25

(c) 73.82 X 60.2

Answer

(a) 95

(b) 2900

(c) 4440

Provide the answer to each addition or subtraction problem with correct significant digits.

(d) 4560 + 12

(e) 11.27 - 5.881

(f) 44.26 + 5.74

Answer

(d) 4570

(e) 5.39

(f) 50.00

Provide the answer to each mixed operations problem with correct significant digits.

(g) 22.82 + (2.21 X 3.04)

(h) (112.52 - 64.388) X 0.23100

(i) 65.70 / (0.755 + 3.145)

Answer

(g) 29.54

(h) 11.12

(i) 16.85

1.2.3: Scientific Notation - Writing Large and Small Numbers

Learning Objectives

- Express a large number or a small number in scientific notation.
- Carry out arithmetical operations and express the final answer in scientific notation

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. In scientific notation, these numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 ($1 \leq N < 10$), and n is a positive or negative integer ($10^0 = 1$). The number 10 is called the base because it is this number that is raised to the power n. Although a base number may have values other than 10, the base number in scientific notation is always 10.

A simple way to convert numbers to scientific notation is to move the decimal point as many places to the left or right as needed to give a number from 1 to 10 (N). The magnitude of n is then determined as follows:

- If the decimal point is moved to the left n places, n is positive.
- If the decimal point is moved to the right n places, n is negative.

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example $1.2.3.1$.

Example 1.2.3.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

a. 637.8
b. 0.0479
c. 7.86
d. 12,378
e. 0.00032
f. 61.06700
g. 2002.080
h. 0.01020

Solution

Solutions to Example 2.2.1

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, n = 2.	6.378×10^2
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, n = −2.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^0 = 1$.)	7.86×10^0
d	Because the decimal point was moved four places to the left, n = 4.	1.2378×10^4
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}

	Explanation	Answer
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^1
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^3
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example $1.2.3.2$illustrates how to do this.

✓ Example 1.2.3.2: Expressing Sums and Differences in Scientific Notation

Carry out the appropriate operation and then express the answer in scientific notation.

a. $(1.36 \times 10^2) + (4.73 \times 10^3)$
b. $(6.923 \times 10^{-3}) - (8.756 \times 10^{-4})$

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) =$ $(1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) =$ $(0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.047

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example $1.2.3.3$

✓ Example 1.2.3.3: Expressing Products and Quotients in Scientific Notation

Perform the appropriate operation and express your answer in scientific notation.

a. $(6.022 \times 10^{23})(6.42 \times 10^{-2})$
b. $\dfrac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$
c. $\dfrac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

Solution to Example 2.2.3

	Explanation	
a	In multiplication, we add the exponents: $(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	3.87×10^{22}
b	In division, we subtract the exponents: $\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}} = \frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^{4}$ In converting 0.183×10^{4} to scientific notation, n has become more negative by 1 because the value of N has increased.	1.83×10^{3}
c	This problem has both multiplication and division: $\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{(8.52 \times 10^{-2})} = \frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$	4.7×10^{-31}

1.2.4: Scientific Dimensional Analysis

Conversion Factors

Many quantities can be expressed in several different ways. The English system measurement of 4 cups is also equal to 2 pints, 1 quart, and $\frac{1}{4}$ of a gallon.

$$4 \text{ cups} = 2 \text{ pints} = 1 \text{ quart} = 0.25 \text{ gallon}$$

Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. This is because the units are different. We can establish the same set of equalities for the metric system:

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 1000 \text{ millimeters}$$

The metric system's use of powers of 10 for all conversions makes this quite simple.

Whenever two quantities are equal, a ratio can be written that is numerically equal to 1. Using the metric examples above:

$$\frac{1 \text{ m}}{100 \text{ cm}} = \frac{100 \text{ cm}}{100 \text{ cm}} = \frac{1 \text{ m}}{1 \text{ m}} = 1$$

The $\frac{1 \text{ m}}{100 \text{ cm}}$ is called a **conversion factor**. A conversion factor is a ratio of equivalent measurements. Because both 1 m and 100 cm represent the exact same length, the value of the conversion factor is 1. The conversion factor is read as "1 meter per 100 centimeters". Other conversion factors from the cup measurement example can be:

$$\frac{4 \text{ cups}}{2 \text{ pints}} = \frac{2 \text{ pints}}{1 \text{ quart}} = \frac{1 \text{ quart}}{0.25 \text{ gallon}} = 1$$

Since the numerator and denominator represent equal quantities in each case, all are valid conversion factors.

Scientific Dimensional Analysis

Conversion factors are used in solving problems in which a certain measurement must be expressed with different units. When a given measurement is multiplied by an appropriate conversion factor, the numerical value changes, but the actual size of the quantity measured remains the same. **Dimensional analysis** is a technique that uses the units (dimensions) of the measurement in order to correctly solve problems. Dimensional analysis is best illustrated with an example.

Example 1.2.4.1

How many seconds are in a day?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 day = 24 hours
- 1 hour = 60 minutes
- 1 minute = 60 seconds

Unknown

The known quantities above represent the conversion factors that we will use. The first conversion factor will have day in the denominator so that the "day" unit will cancel. The second conversion factor will then have hours in the denominator, while the third conversion factor will have minutes in the denominator. As a result, the unit of the last numerator will be seconds and that will be the units for the answer.

Step 2: Calculate.

$$1 \text{ d} \times \frac{24 \text{ hr}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} = 86,400 \text{ s}$$

Applying the first conversion factor, the "d" unit cancels and $1 \times 24 = 24$. Applying the second conversion factor, the "hr" unit cancels and $24 \times 60 = 1440$. Applying the third conversion factor, the "min" unit cancels and $1440 \times 60 = 86,400$. The unit that remains is "s" for seconds.

Step 3: Think about your result.

Seconds is a much smaller unit of time than days, so it makes sense that there are a very large number of seconds in one day.

Summary

- A conversion factor is a ratio of equivalent measurements.
- Dimensional analysis is a technique that uses the units (dimensions) of the measurement to solve problems.

Review

1. What is a conversion factor?
2. What is dimensional analysis?
3. How many meters are in 3.7 km?
4. How many kg in 12980 g?

Explore More

Use the link below to answer the following questions**:** http://www.felderbooks.com/papers/units.html

1. What do we always need to express measurements correctly?
2. What does dimensional analysis tell you?
3. How do you know that you have set the problem up incorrectly?

4. How do you know that you have set the problem up correctly?

1.2.4.1: 2.4.1 Practice Unit Conversions

Goal

Practice more unit conversions (dimensional analysis).

? Exercise 1.2.4.1.1

There are 3.7854 L in a gallon. How many L are in 53.72 gallons?

Answer

203.4 L

? Exercise 1.2.4.1.1

Add exercises text here.

Answer

Add texts here. Do not delete this text first.

Exercise 1.2.4.1.1

Tell whether each number represents an exact number or an uncertain, measured number.

I have 12 coins.

Answer

Exact

The total mass of my coins is 27.420 g.

Answer

Uncertain (measured)

1.2.4.1.1: 2.4.1 Practice Dimensional Analysis

? Exercise 1.2.4.1.1.1

Add exercises text here.

Answer

Add texts here. Do not delete this text first.

Exercise 1.2.4.1.1.1

Tell whether each number represents an exact number or an uncertain, measured number.

I have 12 coins.

Answer

Exact

The total mass of my coins is 27.420 g.

Answer

Uncertain (measured)

1.2.5: Metric Unit Conversions

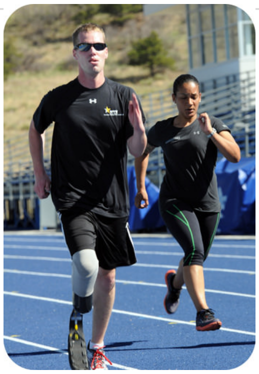


Figure 1.2.5.1 (Credit: Courtesy of the US Army; Source: http://commons.wikimedia.org/wiki/File:Flickr_-_The_U.S._Army_-_Track_practice.jpg(opens in new window); License: Public Domain)

How can a number of track laps be converted to a distance in meters?

You are training for a 10-kilometer run by doing laps on a 400-meter track. You ask yourself how many times you need to run around the track in order to cover ten kilometers (...more than you realize). By using dimensional analysis, you can easily determine the number of laps needed to cover the 10 km distance.

Metric Unit Conversions

The metric system's many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Example 1.2.5.1

A particular experiment for science class requires 120 mL of a solution. The science teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires 120 mL
- $1 \text{ L} = 1000 \text{ mL}$

Unknown

Since each experiment requires 120 mL of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get 4800 mL of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate.

$$4800 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 4.8 \text{ L}$$

Note that the conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3:Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Two-Step Metric Unit Conversions

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (un-prefixed) unit, followed by a conversion to the desired unit. The following example will illustrate this method.

Example 1.2.5.2

Convert $4.3\ \mathrm{cm}$ to $\mu\mathrm{m}$.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1\ \mathrm{m} = 100\ \mathrm{cm}$
- $1\ \mathrm{m} = 10^6\ \mu\mathrm{m}$

Unknown

You may need to consult a table for the multiplication factor represented by each metric prefix. First convert cm to m, then convert m to $\mu\mathrm{m}$.

Step 2: Calculate.

$$4.3\ \mathrm{cm} \times \frac{1\ \mathrm{m}}{100\ \mathrm{cm}} \times \frac{10^6\ \mu\mathrm{m}}{1\ \mathrm{m}} = 43,000\ \mu\mathrm{m}$$

Each conversion factor is written so that the unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

Summary

- Dimensional analysis can be used to carry out metric unit conversions.

Review

1. Perform the following conversions.
    a. 0.074 km to m
    b. 24,600 μg to g
    c. 4.9 × 107 μg to kg
    d. 84 dm to mm

1.2.5.1: Practice Metric Conversions

Exercise 1.2.5.1.1

Convert 0.425 g to kg.

Answer

0.000425 kg

Convert 93.2 mL to L.

Answer

0.0932 L

Convert 8.40 cm to m.

Answer

0.0840 m

Exercise 1.2.5.1.1

Convert 650 Mm to Gm.

Answer

0.65 Gm

Convert 530.0 ng to µg.

Answer

0.5300 µg

Convert 34.5 km to cm.

Answer

3 450 000 cm (3.45 X 10^6 cm)

1.2.6: Density

Learning Objectives

- Define density.
- Use density as a conversion factor.

Density (ρ) is a physical property found by dividing the mass of an object by its volume. Regardless of the sample size, density is always constant. For example, the density of a pure sample of tungsten is always 19.25 grams per cubic centimeter. This means that whether you have one gram or one kilogram of the sample, the density will never vary. The equation, as we already know, is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

or just

$$\rho = \frac{m}{V} \tag{1.2.6.1}$$

Based on this equation, it's clear that density can, and does, vary from element to element and substance to substance due to differences in the relationship of mass and volume. Pure water, for example, has a density of 0.998 g/cm^3 at 25° C. The average densities of some common substances are in Table $1.2.6.1$. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table $1.2.6.1$: Densities of Common Substances

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Density can be measured for all substances—solids, liquids and gases. For solids and liquids, density is often reported using the units of g/cm^3. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

Example $1.2.6.1$: Ethyl Alcohol

Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g

Solution

This is a direct application of Equation $1.2.6.1$:

$$\rho = \frac{23.71002\, g}{30.2\, mL} = 0.785\, g/mL$$

Exercise $1.2.6.1$

a. Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.
b. If you have a 2.130 mL sample of acetic acid with mass 0.002234 kg, what is the density in kg/L?

Answer a

$0.274\ kg/L$

Answer b

$1.049\ kg/L$

Density as a Conversion Factor

Conversion factors can also be constructed for converting between different kinds of units. For example, density can be used to convert between the mass and the volume of a substance. Consider mercury, which is a liquid at room temperature and has a density of 13.6 g/mL. The density tells us that 13.6 g of mercury have a volume of 1 mL. We can write that relationship as follows:

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

$$\frac{13.6\ g}{1\ mL}=1$$

and

$$\frac{1\ mL}{13.6\ g}=1$$

Which one do we use? It depends, as usual, on the units we need to cancel and introduce. For example, suppose we want to know the mass of 2.0 mL of mercury. We would use the conversion factor that has milliliters on the bottom (so that the milliliter unit cancels) and grams on top, so that our final answer has a unit of mass:

$$2.0\ \cancel{\text{mL}}\times\frac{13.6\ \text{g}}{1\ \cancel{\text{mL}}}=27.2\ \text{g}=27\ \text{g}$$

In the last step, we limit our final answer to two significant figures because the volume quantity has only two significant figures; the 1 in the volume unit is considered an exact number, so it does not affect the number of significant figures. The other conversion factor would be useful if we were given a mass and asked to find volume, as the following example illustrates.

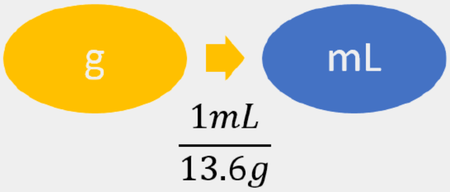
> *Density can be used as a conversion factor between mass and volume.*

✓ Example 1.2.6.2: Mercury Thermometer Steps for Problem Solving

A mercury thermometer for measuring a patient's temperature contains 0.750 g of mercury. What is the volume of this mass of mercury?

Solution

Solution to Example 2.9.2

Steps for Problem Solving	**Unit Conversion**
Identify the "given" information and what the problem is asking you to "find."	Given: 0.750 g Find: mL
List other known quantities.	13.6 g/mL (density of mercury)
Prepare a concept map.	
Calculate.	$0.750\ \cancel{\text{g}}\times\frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}}=0.055147\ldots\ \text{mL}\approx0.0551\ \text{mL}$ We have limited the final answer to three significant figures.

Exercise 1.2.6.2

What is the volume of 100.0 g of air if its density is 1.3 g/L?

Answer

$77\ L$

Summary

- Density is defined as the mass of an object divided by its volume.
- Density can be used as a conversion factor between mass and volume.

Contributions & Attributions

1.2.7: Practice with Density

Exercise 1.2.7.1

What is the density of a material if a sample with a mass of 45.82 g takes up a volume of 8.75 mL?

Answer

5.24 g/mL

What mass would a sample of 32.6 cm^3 of copper be? The density of copper is 8.96 g/cm^3.

Answer

292 g

What volume of acetone would have a mass of 142.5 g? The density of acetone is 0.784 g/mL.

Answer

182 mL

Exercise 1.2.7.1

An empty beaker has a mass of 65.873 g. You add 23.8 mL of solution, and the beaker now has a mass of 85.347 g. What is the density of your solution?

Answer

0.818 g/mL

You have a rectangular block of metal that measures 4.82 cm long, 3.44 cm wide, and 2.65 cm high. The mass is 197.7 g. What is the density of the metal?

Answer

4.50 g/cm^3

To a graduated cylinder with 22.5 mL of water you add 38.95 g of metal pieces. The water level rises to 28.0 mL. What is the density of the metal?

Answer

7.1 g/mL

Exercise 1.2.7.1

You have a copper metal cylinder that is 5.80 cm high. The diameter of the cylinder is 1.44 cm. What would the mass of this cylinder be? The density of copper is 8.96 g/cm^3. The formula to calculate the volume of a cylinder is $V = (\pi) r^2 h$, where r is the radius (half the diameter) and h is the height. If your calculator doesn't have a button for pi, use the value $\pi = 3.14159$.

Hint

The volume of the cylinder is 9.45 cm^3. (9.44589... cm^3 with 3 significant figures)

Answer

84.6 g (If you round the volume to 9.45 cm^3, you get 84.7 g. If you keep the digits in your calculator and round at the very end, it is 84.6 g.)

You have a graduated cylinder with the water level at 30.8 mL. You add 72.88 g of copper to the water. The density of copper is 8.96 g/cm^3. What is the final water level?

Hint

The volume of copper added is 8.13 cm^3, which equals 8.13 mL.

Answer

38.9 mL is the water level after the copper is added.

Before you add acetone to a beaker, the mass is 59.43 g. After you add the acetone, the mass is 81.09 g. The density of acetone is 0.784 g/mL. What volume of acetone is in the beaker?

Hint

The mass of acetone added is 21.66 g.

Answer

27.6 mL

SECTION OVERVIEW

1.3: Matter and Energy

1.3.1: The Classification of Matter

Learning Objectives

- Use physical and chemical properties, including phase, to describe matter.
- Identify a sample of matter as an element, a compound, or a mixture.

Part of understanding matter is being able to describe it. One way chemists describe matter is to assign different kinds of properties to different categories.

Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. **Physical properties** are characteristics that describe matter. They include characteristics such as size, shape, color, and mass. These characteristics can be observed or measured without changing the *identity* of the matter in question. **Chemical properties** are characteristics that describe how matter changes its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. The observation of chemical properties involves a *chemical change* of the matter in question, resulting in matter with a different *identity* and different physical and chemical properties.

Figure 1.3.1.1: (left) Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical as the the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (right) Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

Elements and Compounds

Any sample of matter that has the *same physical and chemical properties throughout* the sample is called a **substance**. There are two types of substances. A substance that cannot be broken down into chemically simpler components is called an **element**. Aluminum, which is used in soda cans and is represented by the symbol Al, is an element. A substance that can be broken down into chemically simpler components (because it consists of more than one element) is called a **compound**. Water is a compound composed of the elements hydrogen and oxygen and is described by the chemical formula, H_2O. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

> Sometimes the word *pure* is used to describe a substance, but this is not absolutely necessary. By definition, any single substance, element or compound is *pure*.

The smallest part of an element that maintains the identity of that element is called an **atom**. Atoms are extremely tiny; to make a line of iron atoms that is 1 inch long, you would need approximately 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a **molecule**. Molecules are composed of two or more atoms that are attached together and behave as a unit. Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the *macroscopic viewpoint* of the universe.

However, scientists can also describe chemical events on the level of individual atoms or molecules, which is referred to as the *microscopic viewpoint*. We will see examples of both macroscopic and microscopic viewpoints throughout this book (Figure 1.3.1.2).

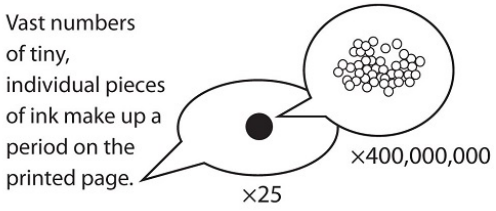


Figure 1.3.1.2: How Many Particles Are Needed for a Period in a Sentence? Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The ×25 and ×400,000,000 indicate the number of times the image is magnified.

Mixtures

A material composed of two or more substances is a **mixture**. In a mixture, the individual substances maintain their chemical identities. Many mixtures are obvious combinations of two or more substances, such as a mixture of sand and water. Such mixtures are called **heterogeneous mixtures**. In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Mixtures with a consistent or uniform composition throughout are called **homogeneous mixtures** (or solutions). For example, when sugar is dissolved in water to form a liquid solution, the individual properties of the components cannot be distinguished. Other examples or homogenous mixtures include solid solutions, like the metal alloy steel, and gaseous solutions, like air which is a mixture of mainly nitrogen and oxygen.

✓ Example 1.3.1.1

How would a chemist categorize each example of matter?

a. saltwater
b. soil
c. water
d. oxygen

Answer a

Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a *homogeneous mixture*, or a solution.

Answer b

Soil is composed of small pieces of a variety of materials, so it is a ***h****eterogeneous mixture*.

Answer c

Water is a *substance*; more specifically, because water is composed of hydrogen and oxygen, it is a *compound*.

Answer d

Oxygen, a *substance*, is an *element*.

? Exercise 1.3.1.2

How would a chemist categorize each example of matter?

a. breakfast coffee
b. hydrogen
c. an egg

Answer a

homogeneous mixture or solution

Answer b

element

Answer c

heterogeneous mixture

Phases or Physical States of Matter

All matter can be further classified by one of three physical **states** or **phases**, solid, liquid or gas. These three descriptions each imply that the matter has certain physical properties when in these states. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers.

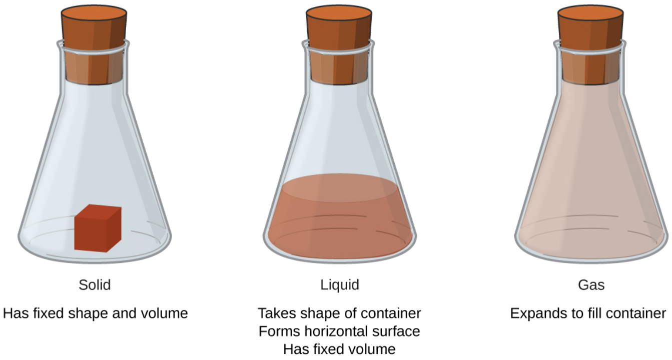


Figure 1.3.1.3: The three most common states or phases of matter are solid, liquid, and gas. (CC BY-4.0; OpenStax)

A beaker labeled solid contains a cube of red matter and says has fixed shape and volume. A beaker labeled liquid contains a brownish-red colored liquid. This beaker says takes shape of container, forms horizontal surfaces, has fixed volume. The beaker labeled gas is filled with a light brown gas. This beaker says expands to fill container.

We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas) (Figure 1.3.1.2).

Figure 1.3.1.5: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. (CC BY-SA 3.0 Unported; Markus Schweiss via Wikipedia)

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a **phase change**, a physical process in which a substance changes from one phase to another (Figure 1.3.1.5). Phase changes are identified by particular names depending on what phases are involved, as summarized in Table 1.3.1.1.

Table 1.3.1.1: Phase Changes

Change	Name
solid to liquid	melting, fusion
solid to gas	sublimation

Change	Name
liquid to gas	boiling, evaporation
liquid to solid	solidification, freezing
gas to liquid	condensation
gas to solid	deposition

Figure 1.3.1.3illustrates the relationships between the different ways matter can be classified.

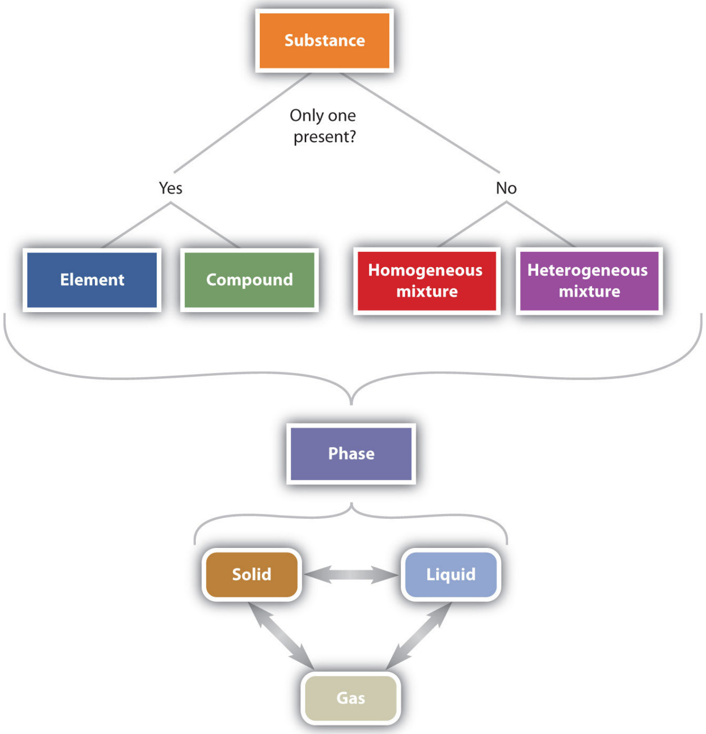


Figure 1.3.1.6: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties.

This table starts with a substance. If there is only one present, it can either be an element or a compound. If there is more than one present, it can either be a homogeneous mixture or a heterogeneous mixture. From these choices, the substance can be in a certain phase. It can either be solid, liquid, or gas, and each phase can go from one to another.

Concept Review Exercises

1. Explain the differences between the physical properties of matter and the chemical properties of matter.
2. What is the difference between a heterogeneous mixture and a homogeneous mixture? Give an example of each.
3. Give at least two examples of a phase change and state the phases involved in each.

Answers

1. Physical properties describe the existence of matter, and chemical properties describe how substances change into other substances.
2. A heterogeneous mixture is obviously a mixture, such as dirt; a homogeneous mixture behaves like a single substance, such as saltwater.
3. solid to liquid (melting) and liquid to gas (boiling) (answers will vary)

Key Takeaways

- Matter can be described with both physical properties and chemical properties.
- Matter can be identified as an element, a compound, or a mixture

? Exercise 1.3.1.3

Does each statement refer to a chemical property or a physical property?

1. Balsa is a very light wood.
2. If held in a flame, magnesium metal burns in air.
3. Mercury has a density of 13.6 g/mL.
4. Human blood is red.

Answer

1. physical property
2. chemical property
3. physical property
4. physical property

? Exercise 1.3.1.4

Does each statement refer to a chemical property or a physical property?

1. The elements sodium and chlorine can combine to make table salt.
2. The metal tungsten does not melt until its temperature exceeds 3,000°C.
3. The ingestion of ethyl alcohol can lead to disorientation and confusion.
4. The boiling point of isopropyl alcohol, which is used to sterilize cuts and scrapes, is lower than the boiling point of water

Answer

1. chemical property
2. physical property
3. chemical property
4. physical property

? Exercise 1.3.1.5

Define *element*. How does it differ from a compound?

Answer

An element is a substance that cannot be broken down into chemically simpler components. Compounds can be broken down into simpler substances.

? Exercise 1.3.1.6

Define *compound*. How does it differ from an element?

Answer

A compound is composed of two or more elements combined in a fixed ratio. An element is the simplest chemical substance.

? Exercise 1.3.1.7

Give two examples of a heterogeneous mixture.

Answer

a salt and pepper mix and a bowl of cereal (answers will vary)

? Exercise 1.3.1.8

Give two examples of a homogeneous mixture.

Answer

vinegar and rubbing alcohol (answers will vary)

? Exercise 1.3.1.9

Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.

1. xenon, a substance that cannot be broken down into chemically simpler components
2. blood, a substance composed of several types of cells suspended in a salty solution called plasma
3. water, a substance composed of hydrogen and oxygen

Answer

1. element
2. heterogeneous mixture
3. compound

? Exercise 1.3.1.10

Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.

1. sugar, a substance composed of carbon, hydrogen, and oxygen
2. hydrogen, the simplest chemical substance
3. dirt, a combination of rocks and decaying plant matter

Answer

1. compound
2. element
3. heterogeneous mixture

? Exercise 1.3.1.11

Identify each substance as an element, a compound, a heterogeneous mixture, or a solution.

1. air, primarily a mixture of nitrogen and oxygen
2. ringer’s lactate, a standard fluid used in medicine that contains salt, potassium, and lactate compounds all dissolved in sterile water
3. tartaric acid, a substance composed of carbon, hydrogen, and oxygen

Answer

1. heterogeneous mixture
2. solution
3. compound

? Exercise 1.3.1.12

What word describes each phase change?

1. solid to liquid
2. liquid to gas
3. solid to gas

Answer

1. melting or fusion
2. boiling or evaporation
3. sublimation

? Exercise 1.3.1.13

1. What word describes each phase change?
    1. liquid to solid
    2. gas to liquid
    3. gas to solid

Answer

1. freezing
2. condensation
3. deposition

1.3.2: Practice Classifying Matter

Exercise 1.3.2.1

Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

a) liquid hand soap (transparent, not creamy)

b) carbonated water (fizzing, not flat)

c) ascorbic acid (also known as vitamin C, $C_6H_8O_6$)

Answer a

homogeneous mixture (has water, various soap molecules, possibly antimicrobial agents)

Answer b

heterogenous mixture (CO_2 gas bubbles are distinct phase, not uniformly mixed)

Answer c

compound (has chemical formula)

Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

d) isopropanol (C_3H_8O)

e) a bottle of 70% isopropanol

f) sulfur

Answer d

compound

Answer e

homogeneous mixture (has 70% isopropanol, 30% water)

Answer f

element

Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

g) bronze (metal alloy of copper and tin)

h) silver

i) oxygen gas (O_2)

Answer g

homogeneous mixture

Answer h

element

Answer i

element (some elements exist in nature as diatomic molecules)

1.3.3: Physical Change

Figure 1.3.3.1 (Credit: User: Paul / Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Vandalized_car.jpg(opens in new window); License: CC by 2.0(opens in new window))

How does a car go from new and shiny to old and beaten up?

Want to buy a car – cheap? Notice there is no specification such as "in good condition" or "needs a little work." The car above is pretty beat up. The body is damaged, the windows are broken, and the interior is probably torn up. But this is still a car. It has all the components of a car, even though you would not want to buy it in the present condition. But change that condition and you have a (possibly) useable car.

Physical Change

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.

Figure 1.3.3.2: Melting ice in the Beaufort Sea. (Credit: Courtesy of Rear Admiral Harley D. Nygren, NOAA; Source: http://commons.wikimedia.org/wiki/File:NOAA_arctic_spring_1950_corp1104.jpg(opens in new window); License: Public Domain)

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust cannot be reconstituted into the same piece of wood that it was before. Cutting grass or pulverizing a rock are examples of irreversible physical changes. Chopping wood for a fire also represents an irreversible physical change, since the pieces cannot be put back together to form the tree.

Figure 1.3.3.3: Chopping wood is an irreversible physical change. (Credit: stevepb; Source: https://pixabay.com/photos/axe-chopper-cut-split-hatchet-674841/(opens in new window); License: Pixabay License)

Physical Changes vs. Chemical Changes

What is the difference between physical and chemical changes? Watch the video to find out.

Summary

- A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not.
- A physical change can be reversible where the original form of the matter can be restored, or irreversible where the original form cannot be restored.
- Melting ice and grinding wood into sawdust are examples of physical changes.

Review

1. Define physical change.
2. Why is melting an ice cube a reversible physical change?
3. Give an example of an irreversible physical change.

1.3.3.1: Practice Classifying Change

Exercise 1.3.3.1.1

Classify each of the following as a chemical or physical change.

a) You let salt water evaporate to leave salt behind.

b) You dilute juice with water because it is too sweet.

c) Starch gets broken down by saliva enzymes into simpler sugars.

Answer a

physical change (liquid water and salt becomes gaseous water and salt, no change in substance identity)

Answer b

physical change (water and juice components becomes more water with juice components, no change in substance identity)

Answer c

chemical change (starch becomes glucose. Try leaving some bread or cracker in your mouth for a while and see if you taste the sweetness!)

Classify each of the following as a chemical or physical change.

d) A solid piece of dry ice becomes gas.

e) You stir some table sugar into water to dissolve it.

f) A piece of lead melts at 327.5 °C.

Answer d

physical change (sublimation)

Answer e

physical change (you still have sugar and water, just mixed)

Answer f

physical change (melting)

Classify each of the following as a chemical or physical change.

g) You drop a piece of zinc metal into acid and it bubbles, tarnishes, and starts to disappear.

h) Apple juice ferments to become hard cider.

i) You add more water to a 15% glucose solution until you have a 10% glucose solution.

Answer g

chemical change (zinc metal is becoming zinc ion and hydrogen gas: zinc ion is soluble in water and hydrogen bubbles away)

Answer h

chemical change (the sugars are being digested, ethanol is being produced: change in substance identity)

Answer i

physical change (still is water and glucose, no change in substance identity)

1.3.4: Heat

Blacksmiths heat solid iron in order to shape it into a variety of different objects. Iron is a rigid, solid metal. At room temperature, it is extremely difficult to bend iron. However, when heated to a high enough temperature, iron can be easily worked. The heat energy in the forge is transferred to the metal, making the iron atoms vibrate more and move around more readily.

Heat is energy that is transferred from one object or substance to another because of a difference in temperature between the two. Heat always flows from an object at a higher temperature to an object at a lower temperature (see figure below). The flow of heat will continue until the two objects are at the same temperature.

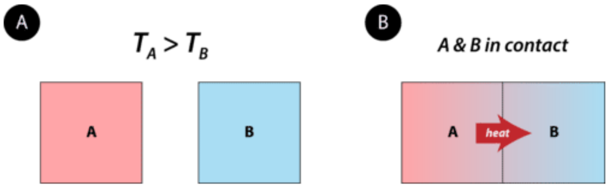


Figure 1.3.4.1: Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.

Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state. When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Summary

- Heat is transferred energy from a site of higher energy to a site of lower energy.
- Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state.

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1.3.4.1: Exothermic and Endothermic Processes

A campfire is an example of basic thermochemistry. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, roast marshmallows, or simply to keep warm when it's cold outside.

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe: *the system* and *the surroundings*. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** is everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings is the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings, and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

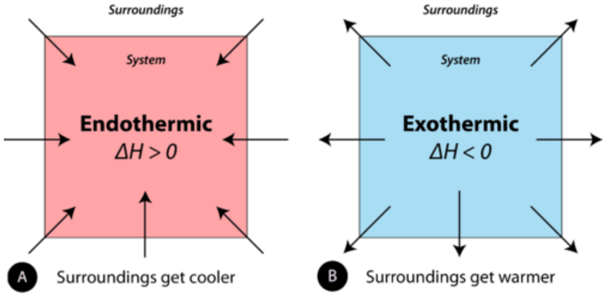

Figure 1.3.4.1.1 : (A) Endothermic reaction. (B) Exothermic reaction. (CC BY-NC; CK-12)

Units of Heat

Heat flow is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by 1^oC. For example, raising the temperature of 100 g of water from 20^oC to 22^oC would require $100 \times 2 = 200 \text{ cal}$.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains 85 kcal or $85,000 \text{ cal}$ In order to make the distinction, the dietary calorie is written with a capital C.

$$1 \text{ kilocalorie} = 1 \text{ Calorie} = 1000 \text{ calories}$$

To say that the snack "contains" 85 Calories means that 85 kcal of energy are released when that snack is processed by the human body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1 \text{ J} = 0.2390 \text{ cal or } 1 \text{ cal} = 4.184 \text{ J}$$

We can calculate the amount of heat released in kilojoules when a 400 Calorie hamburger is digested:

$$400.\ \mathrm{Cal} = 400.\ \mathrm{kcal} \times \frac{4.184\ \mathrm{kJ}}{1\ \mathrm{kcal}} = 1.67 \times 10^3\ \mathrm{kJ}$$

Summary

- The law of conservation of energy states that in any physical or chemical process, energy is neither created nor destroyed.
- A specific portion of matter in a given space that is being studied during an experiment or an observation is the *system*.
- The *surroundings* is everything in the universe that is not part of the system.
- A chemical reaction or physical change is endothermic if heat is absorbed by the system from the surroundings.
- A reaction or change is exothermic if heat is released by the system into the surroundings.

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1.3.4.2: Heat Capacity and Specific Heat

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. The heat capacity of an object depends both on its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the wading pool.

Heat Capacity and Specific Heat

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water under the same sun exposure will not become nearly as hot. This means that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by $1^\circ\mathrm{C}$). Water is very resistant to changes in temperature, while metals generally are not. The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1^\circ\mathrm{C}$. The table below lists the specific heats of some common substances. The symbol for specific heat is c_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree $(\mathrm{J/g^oC})$ or calories per gram per degree $(\mathrm{cal/g^oC})$. This text will use $\mathrm{J/g^oC}$ for specific heat.

Table 1.3.4.2.1 : Specific Heats of Some Common Substances

Substance	Specific Heat $(\mathrm{J/g^oC})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

Notice that water has a very high specific heat compared to most other substances. Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see table above). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

Figure 1.3.4.2.1 : This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake. (CC BY-NC; CK-12)

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by 1°C.
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C.

1.3.4.3: Temperature - Random Motion of Molecules and Atoms

Learning Objectives

- Identify the different between temperature and heat.
- Recognize the different scales used to measure temperature

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.

Figure 1.3.4.3.1 : The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.

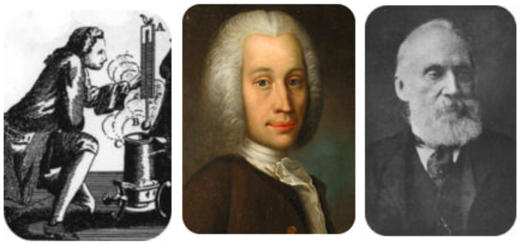

Figure 1.3.4.3.2 : Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32^{o}F$ and the boiling point as $212^{o}F$. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at $0^{o}C$ and $100^{o}C$ respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".

The Kelvin Scale

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of 0 K, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is 273.15 K, while the boiling point is 373.15 K Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales where temperatures are referred to as "degrees F" or "degrees C", we simply designate temperatures in the Kelvin scale as kelvins.

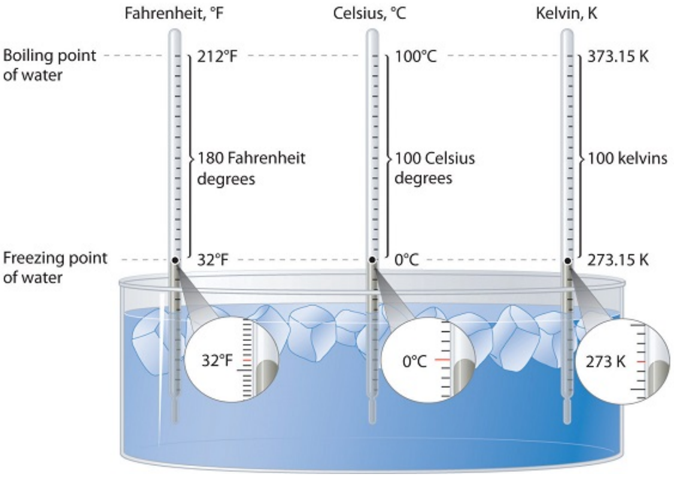

Figure $1.3.4.3.1$: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Converting Between Scales

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in }^\circ\text{C}) + 273.15 = T\,(\text{in K}) \tag{3.10.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in }^\circ\text{C}) \tag{3.10.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales is as follows:

$$^\circ C = \frac{(^\circ F - 32)}{1.8} \tag{3.10.3}$$

$$^\circ F = 1.8 \times (^\circ C) + 32 \tag{3.10.4}$$

✓ Example $1.3.4.3.1$: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

Solution

Converting from Fahrenheit to Celsius requires the use of Equation $3.10.3$:

$$
\begin{aligned}
^\circ C &= \frac{(103.5\,^\circ F - 32)}{1.8} && (1.3.4.3.1)\\
&= 39.7\ ^\circ C && (1.3.4.3.2)
\end{aligned}
$$

Converting from Celsius to Kelvin requires the use of Equation $3.10.1$:

$$K = 39.7\,^{\circ}C + 273.15 \tag{1.3.4.3.3}$$

$$= 312.9\,K \tag{1.3.4.3.4}$$

? Exercise 1.3.4.3.1

Convert each temperature to °C and °F.

a. the temperature of the surface of the sun (5800 K)
b. the boiling point of gold (3080 K)
c. the boiling point of liquid nitrogen (77.36 K)

Answer (a)

5527 K, 9980 °F

Answer (b)

2807 K, 5084 °F

Answer (c)

-195.79 K, -320.42 °F

Summary

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).

Contributions & Attributions

1.3.5: Energy and Heat Capacity Calculations

Learning Objectives

- To relate heat transfer to temperature change.

Heat is a familiar manifestation of transferring energy. When we touch a hot object, energy flows from the hot object into our fingers, and we perceive that incoming energy as the object being “hot.” Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as “cold.” In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Every substance has a characteristic specific heat, which is reported in units of cal/g•°C or cal/g•K, depending on the units used to express Δ*T*. The specific heat of a substance is the amount of energy that must be transferred to or from 1 g of that substance to change its temperature by 1°. Table $1.3.5.1$lists the specific heats for various materials.

Table $1.3.5.1$: *Specific Heats of Some Common Substances*

Substance	**Specific Heat** $(\mathrm{J/g^oC})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

The *direction* of heat flow is not shown in heat = *mc*Δ*T*. If energy goes into an object, the total energy of the object increases, and the values of heat Δ*T* are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and Δ*T* are negative.

Example $1.3.5.1$

A 15.0 g piece of cadmium metal absorbs 134 J of heat while rising from 24.0^oC to 62.7^oC. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Heat} = q = 134 \text{ J}$
- $\text{Mass} = m = 15.0 \text{ g}$
- $\Delta T = 62.7^\circ\text{C} - 24.0^\circ\text{C} = 38.7^\circ\text{C}$

Unknown

- c_p of cadmium $= ? \text{ J/g}^\circ\text{C}$

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^\circ\text{C}} = 0.231 \text{ J/g}^\circ\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known (Table $1.3.5.1$), they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g of water at 23.52°C was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813 \text{ J}}{4.18 \text{ J/g}^\circ\text{C} \times 60.0 \text{ g}} = 3.24^\circ\text{C}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^\circ\text{C} - 3.24^\circ\text{C} = 20.28^\circ\text{C}$$

✓ Example 1.3.5.2

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Solution

We can use heat = $mc\Delta T$ to determine the amount of heat, but first we need to determine ΔT. Because the final temperature of the iron is 73.3°C and the initial temperature is 25.0°C, ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 73.3^\circ\text{C} - 25.0^\circ\text{C} = 48.3^\circ\text{C}$$

The mass is given as 150.0 g, and Table 7.3 gives the specific heat of iron as 0.108 cal/g•°C. Substitute the known values into heat = $mc\Delta T$ and solve for amount of heat:

$$\text{heat} = (150.0 \text{ g}) \left(0.108 \frac{\text{cal}}{\text{g} \cdot {}^\circ\text{C}}\right) (48.3^\circ\text{C}) = 782 \text{ cal}$$

Note how the gram and °C units cancel algebraically, leaving only the calorie unit, which is a unit of heat. Because the temperature of the iron increases, energy (as heat) must be flowing *into* the metal.

? Exercise 1.3.5.1

What quantity of heat is transferred when a 295.5 g block of aluminum metal is cooled from 128.0°C to 22.5°C? What is the direction of heat flow?

Answer

Heat leaves the aluminum block.

✓ Example 1.3.5.2

A 10.3 g sample of a reddish-brown metal gave off 71.7 cal of heat as its temperature decreased from 97.5°C to 22.0°C. What is the specific heat of the metal? Can you identify the metal from the data in Table 1.3.5.1?

Solution

The question gives us the heat, the final and initial temperatures, and the mass of the sample. The value of ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 22.0°\text{C} - 97.5°\text{C} = -75.5°\text{C}$$

If the sample gives off 71.7 cal, it loses energy (as heat), so the value of heat is written as a negative number, −71.7 cal. Substitute the known values into heat = $mc\Delta T$ and solve for c:

$$-71.7 \text{ cal} = (10.3 \text{ g})(c)(-75.5°\text{C})$$

$$c = \frac{-71.7 \text{ cal}}{(10.3 \text{ g})(-75.5°\text{C})}$$

$$c = 0.0923 \text{ cal/g}\bullet°\text{C}$$

This value for specific heat is very close to that given for copper in Table 7.3.

? Exercise 1.3.5.2

A 10.7 g crystal of sodium chloride (NaCl) has an initial temperature of 37.0°C. What is the final temperature of the crystal if 147 cal of heat were supplied to it?

Answer

Summary

Specific heat calculations are illustrated.

1.3.5.1: Specific Heat Calculations

Water has a high capacity for absorbing heat. In a car radiator, it serves to keep the engine cooler than it would otherwise run. As the water circulates through the engine, it absorbs heat from the engine block. When it passes through the radiator, the cooling fan and the exposure to the outside environment allow the water to cool somewhat before it makes another passage through the engine.

Specific Heat Calculations

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Example 1.3.5.1.1

A 15.0 g piece of cadmium metal absorbs 134 J of heat while rising from 24.0^oC to 62.7^oC. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Heat} = q = 134 \text{ J}$
- $\text{Mass} = m = 15.0 \text{ g}$
- $\Delta T = 62.7^\text{o}\text{C} - 24.0^\text{o}\text{C} = 38.7^\text{o}\text{C}$

Unknown

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^\text{o}\text{C}} = 0.231 \text{ J/g}^\text{o}\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known, they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that 60.0 g of water at 23.52^oC was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813 \text{ J}}{4.18 \text{ J/g}^\text{o}\text{C} \times 60.0 \text{ g}} = 3.24^\text{o}\text{C}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^\text{o}\text{C} - 3.24^\text{o}\text{C} = 20.28^\text{o}\text{C}$$

Summary

- The specific heat of a substance can be used to calculate the temperature change of the substance when it is heated or cooled.
- Specific heat calculations are illustrated.

1.3.5.2: Practice with Heat and Temperature Calculations

Exercise 1.3.5.2.1

Convert each Celsius temperature to Fahrenheit and to Kelvin. Use correct significant digits.

a) 62.7 °C

b) 4.8 °C

c) – 51.7 °C

Answer a

145 °F, 335.9 K

Answer b

40.6 °F, 278.0 K

Answer c

– 61.1 °F, 221.5 K

Convert each Fahrenheit temperature to Celsius and to Kelvin. Use correct significant digits.

d) – 46.7 °F

e) 6.3 °F

f) 65 °F

Answer d

– 43.7 °C, 229.4 K

Answer e

– 14.3 °C, 258.9 K

Answer f

18 °C, 291 K

Exercise 1.3.5.2.1

Convert each energy below. [1 cal = 4.184 J exactly] [1 Cal = 1 kcal = 10^3 cal]

a) 4.83 cal to J and to Cal

b) 657 J to cal and to kcal

c) 54.0 Cal to cal and to J

Answer a

20.2 J, 0.00483 Cal

Answer b

157 cal, 0.157 kcal

Answer c

54000 cal (or 5.40 x 10^4 cal), 226000 J. (3 sig dig in each)

Exercise 1.3.5.2.1

You find that it takes 261 J to raise the temperature of a piece of metal from 18.6 °C to 42.2 °C. The mass of your metal is 28.72 g. What is the specific heat of the metal?

Answer

0.385 J/g °C

Exercise 1.3.5.2.1

How much energy is needed to raise the temperature of 65.8 g of aluminum from 22.0 °C to 47.8 °C? The specific heat of aluminum is 0.214 cal/g°C.

Answer

363 cal (which is 1520 J)

1.3.6: Phase Changes

Learning Objectives

- Determine the heat associated with a phase change.

Matter can exist in one of several different states, including a gas, liquid, or solid state. The amount of energy in molecules of matter determines the **state of matter**.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

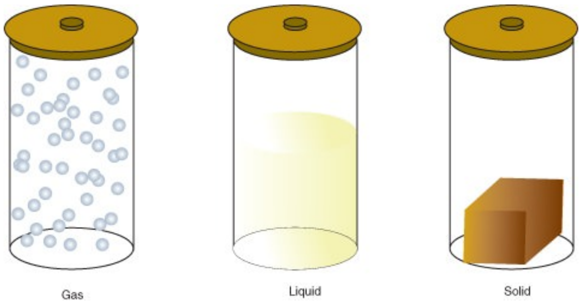

Figure 1.3.6.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

The following are the changes of state:

Changes of State

Solid → Liquid	Melting or fusion
Liquid → Gas	Vaporization
Liquid → Solid	Freezing
Gas → Liquid	Condensation
Solid → Gas	Sublimation

- If heat is added to a substance, such as in melting, vaporization, and sublimation, the process is **endothermic**. In this instance, heat is increasing the speed of the molecules causing them move faster (examples: solid to liquid; liquid to gas; solid to gas).
- If heat is removed from a substance, such as in freezing and condensation, then the process is **exothermic**. In this instance, heat is decreasing the speed of the molecules causing them move slower (examples: liquid to solid; gas to liquid). These changes **release heat** to the surroundings.
- The amount of heat needed to change a sample from solid to liquid would be the same to reverse from liquid to solid. The only difference is the direction of heat transfer.

Example 1.3.6.1

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. ice forming on a pond

Solution

a. endothermic - you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b. exothermic - think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

ise

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer

a. exothermic

b. endothermic

A phase change is a physical process in which a substance goes from one phase to another. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance. The melting point is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid). The boiling point is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid). The nature of the phase change depends on the direction of the heat transfer. Heat going *into* a substance changes it from a solid to a liquid or a liquid to a gas. Removing heat *from* a substance changes a gas to a liquid or a liquid to a solid.

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O. If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O. A similar process can occur at 100°C: adding heat increases the amount of gaseous H_2O, while removing heat increases the amount of liquid H_2O (Figure $1.3.6.1$).

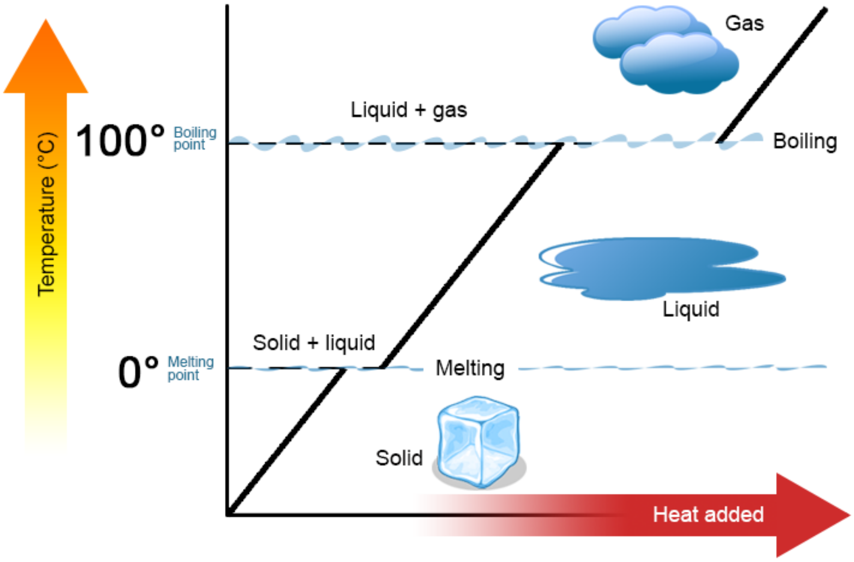

Figure $1.3.6.2$: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).

> Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well.

Second, as shown in Figure 1.3.6.1, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means “constant temperature”). Again, consider H_2O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H_2O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

For each phase change of a substance, there is a characteristic quantity of heat needed to perform the phase change per gram (or per mole) of material. The heat of fusion (ΔH_{fus}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the melting point. The heat of vaporization (ΔH_{vap}) is the amount of heat per gram (or per mole) required for a phase change that occurs at the boiling point. If you know the total number of grams or moles of material, you can use the ΔH_{fus} or the ΔH_{vap} to determine the total heat being transferred for melting or solidification using these expressions:

$$\text{heat} = n \times \Delta H_{fus} \tag{1.3.6.1}$$

where n is the number of moles and ΔH_{fus} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{fus} \tag{1.3.6.2}$$

where m is the mass in grams and ΔH_{fus} is expressed in energy/gram.

For the boiling or condensation, use these expressions:

$$\text{heat} = n \times \Delta H_{vap} \tag{1.3.6.3}$$

where n is the number of moles) and ΔH_{vap} is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{vap} \tag{1.3.6.4}$$

where m is the mass in grams and ΔH_{vap} is expressed in energy/gram.

Remember that a phase change depends on the direction of the heat transfer. If heat transfers in, solids become liquids, and liquids become solids at the melting and boiling points, respectively. If heat transfers out, liquids solidify, and gases condense into liquids. At these points, there are no changes in temperature as reflected in the above equations.

✓ Example 1.3.6.2

How much heat is necessary to melt 55.8 g of ice (solid H_2O) at 0°C? The heat of fusion of H_2O is 79.9 cal/g.

Solution

We can use the relationship between heat and the heat of fusion (Equation 1.3.6.1) to determine how many cal of heat are needed to melt this ice:

$$\begin{aligned} \text{heat } &= \mathrm{m} \times \Delta \mathrm{H}_{\text{fus}} \\ \text{heat } &= (55.8 \ \cancel{\mathrm{g}}) \left(\frac{79.9 \text{ cal}}{\cancel{\mathrm{g}}} \right) = 4,460 \text{ cal} \end{aligned}$$

? Exercise 1.3.6.2

How much heat is necessary to vaporize 685 g of H_2O at 100°C? The heat of vaporization of H_2O is 540 cal/g.

Answer

$$\begin{aligned} \text{heat } &= \mathrm{m} \times \Delta \mathrm{H}_{\text{vap}} \\ \text{heat } &= (685 \ \cancel{\mathrm{g}}) \left(\frac{540 \text{ cal}}{\cancel{\mathrm{g}}} \right) = 370,000 \text{ cal} \end{aligned}$$

Table 1.3.6.1 lists the heats of fusion and vaporization for some common substances. Note the units on these quantities; when you use these values in problem solving, make sure that the other variables in your calculation are expressed in units consistent with the units in the specific heats or the heats of fusion and vaporization.

Table 1.3.6.1: Heats of Fusion and Vaporization for Selected Substances

Substance	ΔH_{fus} (cal/g)	ΔH_{vap} (cal/g)
aluminum (Al)	94.0	2,602
gold (Au)	15.3	409
iron (Fe)	63.2	1,504
water (H_2O)	79.9	540
sodium chloride (NaCl)	123.5	691
ethanol (C_2H_5OH)	45.2	200.3
benzene (C_6H_6)	30.4	94.1

Sublimation

There is also a phase change where a solid goes directly to a gas:

$$\text{solid} \rightarrow \text{gas} \tag{1.3.6.5}$$

This phase change is called *sublimation*. Each substance has a characteristic heat of sublimation associated with this process. For example, the heat of sublimation (ΔH_{sub}) of H_2O is 620 cal/g.

We encounter sublimation in several ways. You may already be familiar with dry ice, which is simply solid carbon dioxide (CO_2). At −78.5°C (−109°F), solid carbon dioxide sublimes, changing directly from the solid phase to the gas phase:

$$CO_2(s) \xrightarrow{-78.5^\circ C} CO_2(g) \tag{1.3.6.6}$$

Solid carbon dioxide is called dry ice because it does not pass through the liquid phase. Instead, it does directly to the gas phase. (Carbon dioxide *can* exist as liquid but only under high pressure.) Dry ice has many practical uses, including the long-term preservation of medical samples.

Even at temperatures below 0°C, solid H_2O will slowly sublime. For example, a thin layer of snow or frost on the ground may slowly disappear as the solid H_2O sublimes, even though the outside temperature may be below the freezing point of water. Similarly, ice cubes in a freezer may get smaller over time. Although frozen, the solid water slowly sublimes, redepositing on the colder cooling elements of the freezer, which necessitates periodic defrosting (frost-free freezers minimize this redeposition). Lowering the temperature in a freezer will reduce the need to defrost as often.

Under similar circumstances, water will also sublime from frozen foods (e.g., meats or vegetables), giving them an unattractive, mottled appearance called freezer burn. It is not really a “burn,” and the food has not necessarily gone bad, although it looks unappetizing. Freezer burn can be minimized by lowering a freezer’s temperature and by wrapping foods tightly so water does not have any space to sublime into.

Key Takeaway

- There is an energy change associated with any phase change.

1.3.6.1: Practice Heat to Melt or Boil

Exercise 1.3.6.1.1

How much heat does it take to melt 43.77 g of copper if the copper is already at its melting point? The heat of fusion of copper is 49 cal/g.

Answer

2100 cal (2144.73 rounded to 2 sig. dig.)

Exercise 1.3.6.1.1

How much heat does it take to boil 26.88 g of ethanol if the ethanol is already at its boiling point? The heat of vaporization of ethanol is 853 J/g.

Answer

22900 J

Exercise 1.3.6.1.1

What if you start with 85.5 g of ethanol that is only 18.5 °C? The boiling point of ethanol is 78.4 °C. The heat of vaporization of ethanol is 853 J/g. The specific heat of ethanol is 2.46 J/g°C.

Answer

85500 J (It takes 12600 J to heat it from 18.5°C to 78.4°C plus another 72900 J to boil it.)

Exercise 1.3.6.1.1

Starting with 245 g of ice that is –16.3°C, how much energy is needed to warm the ice, melt it, and warm the water to 27.2 °C?

The specific heat of ice is 0.488 cal/g°C, the specific heat of liquid water is 1.00 cal/g°C, and the heat of fusion of ice is 79.7 cal/g.

Answer

28100 cal (It takes 1950 cal to warm the ice to 0.0°C, plus 19500 cal to melt the ice, plus 6660 cal to warm the water to 27.2°C).

SECTION OVERVIEW

1.4: Atoms and Elements

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1.4.1: Cutting Alumimun until you get Atoms

Take some aluminum foil. Cut it in half. Now there are two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil. It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can this exercise be taken, at least in theory? Can one continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC by Democritus and other ancient Greek philosophers (Figure 1.4.1.1).

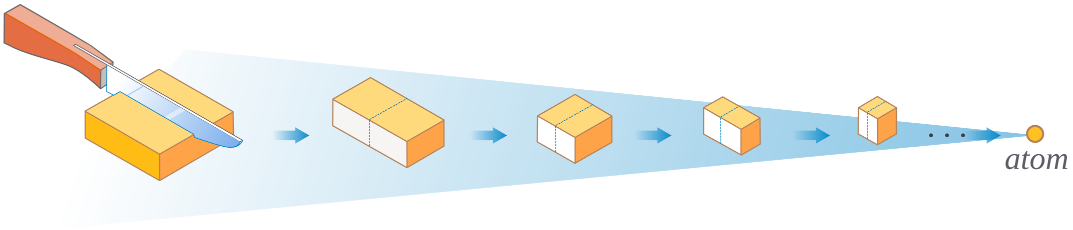


Figure 1.4.1.1: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya via LibreTexts)

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called diatomic molecules. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

1.4.1.1: Indivisible - The Atomic Theory

Learning Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

You learned earlier that all matter in the universe is made out of tiny building blocks called atoms. All modern scientists accept the concept of the atom, but when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus and the Greek Philosophers

About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge, entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable "thing." One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon—tiny, indivisible, solid objects—make up all matter in the universe.

Figure 1.4.1.1.1 : (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

Democritus then reasoned that changes occur when the many atomos in an object were reconnected or recombined in different ways. Democritus even extended this theory, suggesting that there were different varieties of atoms with different shapes, sizes, and masses. He thought, however, that shape, size, and mass were the only properties differentiating the different types of atomos. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves, but rather, resulted from the different ways in which the atomos were combined and connected to one another.

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they are known today) was fully appreciated.

Greeks: "All Thought and No Action"

Greek philosophers were "all thought and no action" and did not feel the need to test their theories with reality. In contrast, Dalton's efforts were based on experimentation and testing ideas against reality.

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

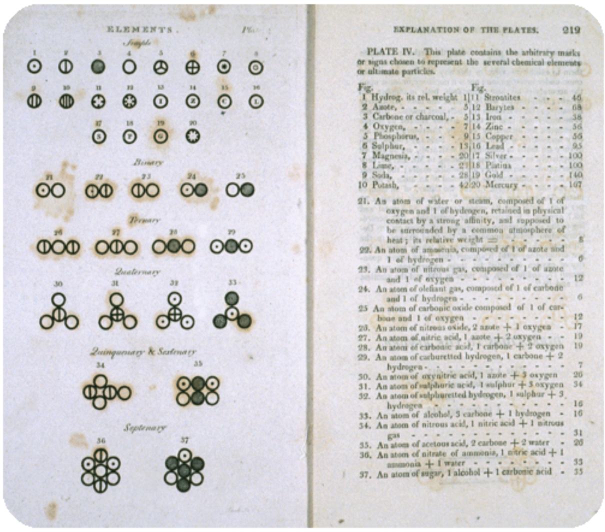
The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 1.4.1.1.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory were as follows:

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.



EXPLANATION OF THE PLATES. 219

PLATE IV. This plate contains the arbitrary marks or signs chosen to represent the several chemical elements or ultimate particles.

Fig.			Fig.		
1	Hydrog. its rel. weight	1	11	Stroutites	46
2	Azote,	5	12	Barytes	68
3	Carbone or charcoal,	5	13	Iron	38
4	Oxygen,	7	14	Zinc	56
5	Phosphorus,	9	15	Copper	56
6	Sulphur,	13	16	Lead	95
7	Magnesia,	20	17	Silver	100
8	Lime,	24	18	Platina	100
9	Soda,	28	19	Gold	140
10	Potash,	42	20	Mercury	167

21. An atom of water or steam, composed of 1 of oxygen and 1 of hydrogen, retained in physical contact by a strong affinity, and supposed to be surrounded by a common atmosphere of heat; its relative weight = 8
22. An atom of ammonia, composed of 1 of azote and 1 of hydrogen 6
23. An atom of nitrous gas, composed of 1 of azote and 1 of oxygen 12
24. An atom of olefiant gas, composed of 1 of carbone and 1 of hydrogen 6
25 An atom of carbonic oxide composed of 1 of carbone and 1 of oxygen 12
26. An atom of nitrous oxide, 2 azote + 1 oxygen 17
27. An atom of nitric acid, 1 azote + 2 oxygen 19
28. An atom of carbonic acid, 1 carbone + 2 oxygen 19
29. An atom of carburetted hydrogen, 1 carbone + 2 hydrogen 7
30. An atom of oxynitric acid, 1 azote + 3 oxygen 26
31. An atom of sulphuric acid, 1 sulphur + 3 oxygen 34
32. An atom of sulphuretted hydrogen, 1 sulphur + 3 hydrogen 16
33. An atom of alcohol, 3 carbone + 1 hydrogen 16
34. An atom of nitrous acid, 1 nitric acid + 1 nitrous gas 31
35. An atom of acetous acid, 2 carbone + 2 water 26
36. An atom of nitrate of ammonia, 1 nitric acid + 1 ammonia + 1 water 33
37. An atom of sugar, 1 alcohol + 1 carbonic acid 35

Figure 1.4.1.1.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video 1.4.1.1.1, control of individual atoms can be use used create animations.

Video 1.4.1.1.1 : A Boy And His Atom: The World's Smallest Movie.

A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos." However, other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.
- Dalton's Atomic Theory is the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom. The general tenets of this theory are:
    - All matter is composed of extremely small particles called atoms.
    - Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
    - Atoms cannot be subdivided, created, or destroyed.
    - Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
    - In chemical reactions, atoms are combined, separated, or rearranged.

1.4.2: The Nuclear Atom

Learning Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" model of the atom and the evidence for it.
- Draw a diagram of Thomson's "plum pudding" model of the atom and explain why it has this name.
- Describe Rutherford's gold foil experiment and explain how this experiment altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.

Figure 1.4.2.1: J. J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom.

Electrons and Plums

The electron was discovered by J. J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This was a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or similar to a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges, and the electron would be pulled back more toward the center region of the atom.

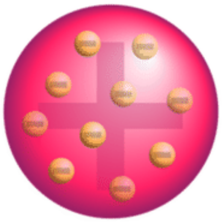


Figure 1.4.2.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871-1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Atoms and Gold

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

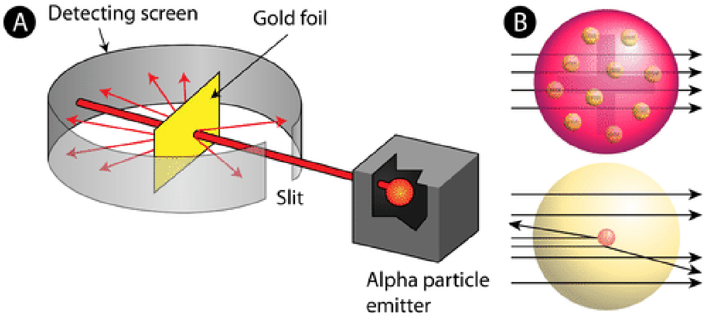


Figure 1.4.2.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupy the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Atomic Nucleus

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. The figure below shows these parts of the atom.

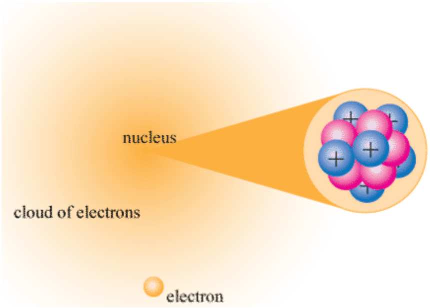


Figure 1.4.2.4: The nuclear atom.

The nucleus of the atom is extremely small. Its radius is only about 1/100,000 of the total radius of the atom. Electrons have virtually no mass, but protons and neutrons have a lot of mass for their size. As a result, the nucleus has virtually all the mass of an atom. Given its great mass and tiny size, the nucleus is very dense. If an object the size of a penny had the same density as the nucleus of an atom, its mass would be greater than 30 million tons!

Holding it all Together

Particles with opposite electric charges attract each other. This explains why negative electrons orbit the positive nucleus. Particles with the same electric charge repel each other. This means that the positive protons in the nucleus push apart from one another. So why doesn't the nucleus fly apart? An even stronger force—called the **strong nuclear force**—holds protons and neutrons together in the nucleus.

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

1.4.2.1: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it was not entirely accurate, because contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to us: protons and neutrons. We already learned that J. J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus, called a **proton**. There is also a third subatomic particle, known as a neutron.

Electrons

Electrons are one of three main types of particles that make up atoms. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particle called leptons. All leptons have an electric charge of -1 or 0. Electrons are extremely small. The mass of an electron is only about 1/2000 the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of -1, which is equal but opposite to the charge of a proton, which is $+1$. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract one another, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. However, this is an incorrect perspective, as quantum mechanics demonstrates that electrons are more complicated.

Figure 1.4.2.1.1 : Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

A proton is one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one $(+1)$ and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom.

Neutrons

Atoms of all elements—except for most atoms of hydrogen—have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge—they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they are bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 1.4.2.1.1 : Properties of Subatomic Particles

Particle	**Symbol**	**Mass (amu)**	**Relative Mass (proton = 1)**	**Relative Charge**	**Location**
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	−1	outside the nucleus
neutron	n^0	1	1	0	inside the nucleus

Table $1.4.2.1.1$ gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they are neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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1.4.2.2: Isotopes - When the Number of Neutrons Varies

Learning Objectives

- Explain what isotopes are and how an isotope affects an element's atomic mass.
- Determine the number of protons, electrons, and neutrons of an element with a given mass number.

All atoms of the same element have the same number of protons, but some may have different numbers of neutrons. For example, all carbon atoms have six protons, and most have six neutrons as well. But some carbon atoms have seven or eight neutrons instead of the usual six. Atoms of the same element that differ in their numbers of neutrons are called **isotopes**. Many isotopes occur naturally. Usually one or two isotopes of an element are the most stable and common. Different isotopes of an element generally have the same physical and chemical properties because they have the same numbers of protons and electrons.

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure $1.4.2.2.1$. Most hydrogen atoms have just one proton, one electron, and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

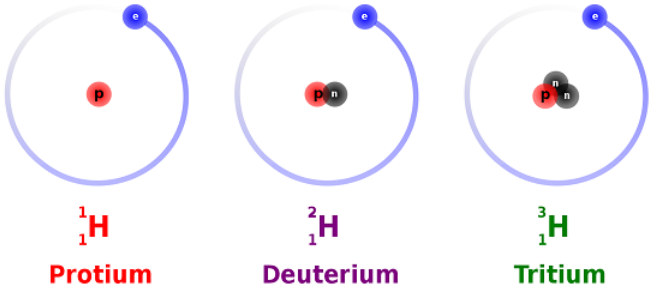


Figure $1.4.2.2.1$: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

For most elements other than hydrogen, isotopes are named for their mass number. For example, carbon atoms with the usual 6 neutrons have a mass number of 12 (6 protons + 6 neutrons = 12), so they are called carbon-12. Carbon atoms with 7 neutrons have an atomic mass of 13 (6 protons + 7 neutrons = 13). These atoms are the isotope called carbon-13.

Example $1.4.2.2.1$: Lithium Isotopes

a. What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons?
b. What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons?

Solution

A lithium atom contains 3 protons in its nucleus irrespective of the number of neutrons or electrons.

a.

$$\begin{aligned} \text{atomic number} = (\text{number of protons}) &= 3 \\ (\text{number of neutrons}) &= 3 \end{aligned}$$

$$\begin{aligned} \text{mass number } &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number } &= 3 + 3 \\ &= 6 \end{aligned}$$

b.

$$\begin{aligned} \text{atomic number} = (\text{number of protons}) &= 3 \\ (\text{number of neutrons}) &= 4 \end{aligned}$$

$$\begin{aligned}\text{mass number } &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number } &= 3 + 4 \\ &= 7\end{aligned}$$

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons, and as an isotope with 4 neutrons, but it doesn't exist as an isotope with 2 neutrons or as an isotope with 5 neutrons.

Stability of Isotopes

Atoms need a certain ratio of neutrons to protons to have a stable nucleus. Having too many or too few neutrons relative to protons results in an unstable, or radioactive, nucleus that will sooner or later break down to a more stable form. This process is called radioactive decay. Many isotopes have radioactive nuclei, and these isotopes are referred to as radioisotopes. When they decay, they release particles that may be harmful. This is why radioactive isotopes are dangerous and why working with them requires special suits for protection. The isotope of carbon known as carbon-14 is an example of a radioisotope. In contrast, the carbon isotopes called carbon-12 and carbon-13 are stable.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundance"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element—chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below:

Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

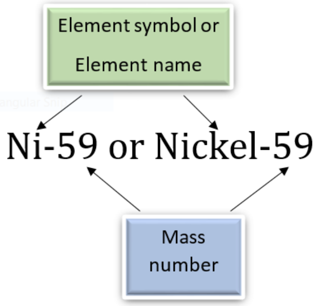
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\text{Ni}$$

$$^{238}_{92}\text{U}$$

In the nickel nucleus represented above, the atomic number 28 indicates that the nucleus contains 28 protons, and therefore, it must contain 31 neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons, as all uranium nuclei do; and this particular uranium nucleus has 146 neutrons.

Another way of representing isotopes is by adding a hyphen and the mass number to the chemical name or symbol. Thus the two nuclei would be Nickel-59 or Ni-59 and Uranium-238 or U-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) are given to the side of the name.



✓ Example 1.4.2.2.2: Potassium-40

How many protons, electrons, and neutrons are in an atom of ${}^{40}_{19}\text{K}$?

Solution

$$\text{atomic number} = (\text{number of protons}) = 19$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 19$$

The mass number, 40, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 40 - 19 = 21.$$

✓ Example 1.4.2.2.3: Zinc-65

How many protons, electrons, and neutrons are in an atom of zinc-65?

Solution

$$\text{number of protons} = 30$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 30$$

The mass number, 65, is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 65 - 30 = 35$$

? Exercise 1.4.2.2.3

How many protons, electrons, and neutrons are in each atom?

a. ${}^{60}_{27}\text{Co}$
b. Na-24
c. ${}^{45}_{20}\text{Ca}$
d. Sr-90

Answer a:
27 protons, 27 electrons, 33 neutrons

Answer b:
11 protons, 11 electrons, 13 neutrons

Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Summary

- The number of protons is always the same in atoms of the same element.
- The number of neutrons can be different, even in atoms of the same element.
- Atoms of the same element that contain the same number of protons, but different numbers of neutrons, are known as **isotopes**.
- Isotopes of any given element all contain the same number of protons, so they have the same atomic number (for example, the atomic number of helium is always 2).
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

1.4.2.3: Elements- Defined by Their Number of Protons

Learning Objectives

- Define atomic number.
- Define mass number.
- Determine the number of protons, neutrons, and electrons in an atom.

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as H for hydrogen, C for carbon, or He for helium. However, it would more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.

Figure 1.4.2.3.1: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 1.4.2.3.1). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (Z). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 1.4.2.3.1: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, 79 electrons as well.

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table 1.4.2.3.1). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table 1.4.2.3.1).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units ($2\ \mathrm{amu}$ for the protons + $2\ \mathrm{amu}$ for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called isotopes and are discussed in detail later on).

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

An atom's mass number is very easy to calculate, provided that you know the number of protons and neutrons in an atom.

✓ Example 4.5.1

What is the mass number of an atom of helium that contains 2 neutrons?

Solution

$(\text{number of protons}) = 2$ (Remember that an atom of helium always has 2 protons.)

$(\text{number of neutrons}) = 2$

$\text{mass number} = (\text{number of protons}) + (\text{number of neutrons})$

$$\text{mass number} = \mathbf{2 + 2 = 4}$$

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are O for oxygen, Zn for zinc, and Fe for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table 1.4.2.3.2

Table 1.4.2.3.2 : *Symbols and Latin Names for Elements*

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum
Sn	Tin	Stannum

Chemical Symbol	Name	Latin Name
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

1.4.2.4: Practice Isotopes

Exercise 1.4.2.4.1

For each isotope shown, give the number of protons, neutrons, and electrons. These isotopes are neutral (charge = 0).

a) ${}^{11}{}_{5}B$

b) ${}^{41}{}_{19}K$

c) ${}^{70}{}_{32}Ge$

Answer a

5 protons, 6 neutrons, 5 electrons

Answer b

19 protons, 22 neutrons, 19 electrons

Answer c

32 protons, 38 neutrons, 32 electrons

Exercise 1.4.2.4.1

For each isotope shown, give the number of protons, neutrons, and electrons.

d) ${}^{44}{}_{20}Ca^{2+}$

e) ${}^{41}{}_{19}K^{1+}$

f) ${}^{18}{}_{8}O^{2-}$

Answer d

20 protons, 24 neutrons, 18 electrons

Answer e

19 protons, 22 neutrons, 18 electrons

Answer f

8 protons, 10 neutrons, 10 electrons

Exercise 1.4.2.4.1

Given the number of protons, neutrons, and electrons, what would be the mass number and the charge number of each isotope? What would be the atomic number? The identity of the element?

g) 7 protons, 7 neutrons, 10 electrons

h) 15 protons, 16 neutrons, 15 electrons

i) 25 protons, 30 neutrons, 23 electrons

Answer g

Mass number is 14, charge is –3. Atomic number is 7, identity is nitrogen (N).

Answer h

Mass number is 31, charge is 0. Atomic number is 15, identity is phosphorus (p).

Answer i

Mass number is 55, charge is +2. Atomic number is 25, identity is manganese (Mn).

1.4.3: Looking for Patterns - The Periodic Table

Learning Objectives

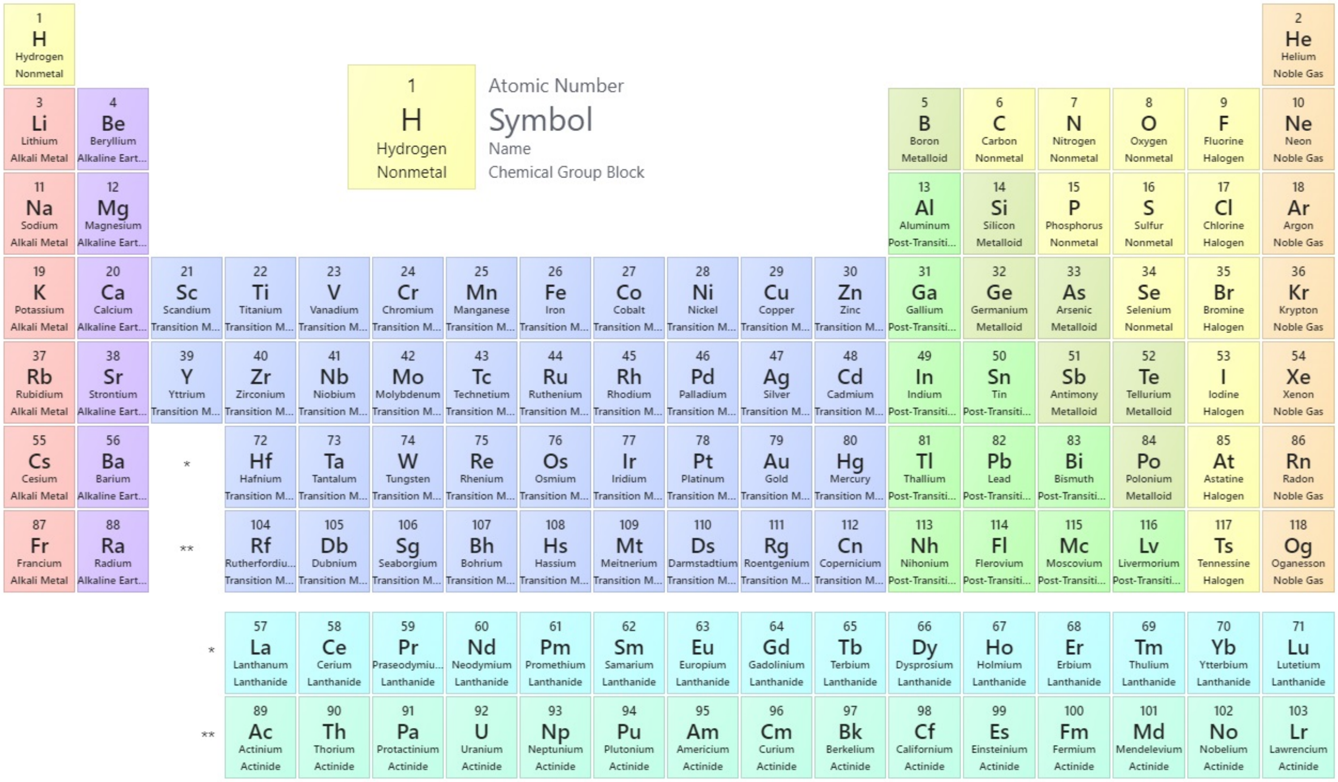
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

> *Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.*

The periodic table is one of the cornerstones of chemistry because it organizes all of the known elements on the basis of their chemical properties. A modern version is shown in Figure 1.4.3.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.



1 H Hydrogen Nonmetal	Atomic Number Symbol Name Chemical Group Block

1 H Hydrogen Nonmetal																	2 He Helium Noble Gas
3 Li Lithium Alkali Metal	4 Be Beryllium Alkaline Eart...											5 B Boron Metalloid	6 C Carbon Nonmetal	7 N Nitrogen Nonmetal	8 O Oxygen Nonmetal	9 F Fluorine Halogen	10 Ne Neon Noble Gas
11 Na Sodium Alkali Metal	12 Mg Magnesium Alkaline Eart...											13 Al Aluminum Post-Transiti...	14 Si Silicon Metalloid	15 P Phosphorus Nonmetal	16 S Sulfur Nonmetal	17 Cl Chlorine Halogen	18 Ar Argon Noble Gas
19 K Potassium Alkali Metal	20 Ca Calcium Alkaline Eart...	21 Sc Scandium Transition M...	22 Ti Titanium Transition M...	23 V Vanadium Transition M...	24 Cr Chromium Transition M...	25 Mn Manganese Transition M...	26 Fe Iron Transition M...	27 Co Cobalt Transition M...	28 Ni Nickel Transition M...	29 Cu Copper Transition M...	30 Zn Zinc Transition M...	31 Ga Gallium Post-Transiti...	32 Ge Germanium Metalloid	33 As Arsenic Metalloid	34 Se Selenium Nonmetal	35 Br Bromine Halogen	36 Kr Krypton Noble Gas
37 Rb Rubidium Alkali Metal	38 Sr Strontium Alkaline Eart...	39 Y Yttrium Transition M...	40 Zr Zirconium Transition M...	41 Nb Niobium Transition M...	42 Mo Molybdenum Transition M...	43 Tc Technetium Transition M...	44 Ru Ruthenium Transition M...	45 Rh Rhodium Transition M...	46 Pd Palladium Transition M...	47 Ag Silver Transition M...	48 Cd Cadmium Transition M...	49 In Indium Post-Transiti...	50 Sn Tin Post-Transiti...	51 Sb Antimony Metalloid	52 Te Tellurium Metalloid	53 I Iodine Halogen	54 Xe Xenon Noble Gas
55 Cs Cesium Alkali Metal	56 Ba Barium Alkaline Eart...	*	72 Hf Hafnium Transition M...	73 Ta Tantalum Transition M...	74 W Tungsten Transition M...	75 Re Rhenium Transition M...	76 Os Osmium Transition M...	77 Ir Iridium Transition M...	78 Pt Platinum Transition M...	79 Au Gold Transition M...	80 Hg Mercury Transition M...	81 Tl Thallium Post-Transiti...	82 Pb Lead Post-Transiti...	83 Bi Bismuth Post-Transiti...	84 Po Polonium Metalloid	85 At Astatine Halogen	86 Rn Radon Noble Gas
87 Fr Francium Alkali Metal	88 Ra Radium Alkaline Eart...	**	104 Rf Rutherfordiu... Transition M...	105 Db Dubnium Transition M...	106 Sg Seaborgium Transition M...	107 Bh Bohrium Transition M...	108 Hs Hassium Transition M...	109 Mt Meitnerium Transition M...	110 Ds Darmstadtium Transition M...	111 Rg Roentgenium Transition M...	112 Cn Copernicium Transition M...	113 Nh Nihonium Post-Transiti...	114 Fl Flerovium Post-Transiti...	115 Mc Moscovium Post-Transiti...	116 Lv Livermorium Post-Transiti...	117 Ts Tennessine Halogen	118 Og Oganesson Noble Gas
		*	57 La Lanthanum Lanthanide	58 Ce Cerium Lanthanide	59 Pr Praseodymiu... Lanthanide	60 Nd Neodymium Lanthanide	61 Pm Promethium Lanthanide	62 Sm Samarium Lanthanide	63 Eu Europium Lanthanide	64 Gd Gadolinium Lanthanide	65 Tb Terbium Lanthanide	66 Dy Dysprosium Lanthanide	67 Ho Holmium Lanthanide	68 Er Erbium Lanthanide	69 Tm Thulium Lanthanide	70 Yb Ytterbium Lanthanide	71 Lu Lutetium Lanthanide
		**	89 Ac Actinium Actinide	90 Th Thorium Actinide	91 Pa Protactinium Actinide	92 U Uranium Actinide	93 Np Neptunium Actinide	94 Pu Plutonium Actinide	95 Am Americium Actinide	96 Cm Curium Actinide	97 Bk Berkelium Actinide	98 Cf Californium Actinide	99 Es Einsteinium Actinide	100 Fm Fermium Actinide	101 Md Mendelevium Actinide	102 No Nobelium Actinide	103 Lr Lawrencium Actinide

Figure 1.4.3.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

> *The word halogen comes from the Greek for "salt maker" because these elements combine with other elements to form a group of compounds called salts.*

To Your Health: Radon

Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building's inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

Each row of elements on the periodic table is called a period. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 1.4.3.2 A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 1.4.3.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

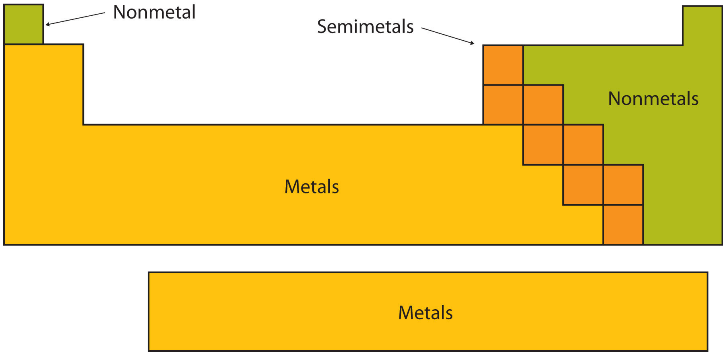


Figure 1.4.3.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

✓ Example 1.4.3.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

a. Se

b. Mg
c. Ge

Solution

a. In Figure 1.4.3.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

? Exercise 1.4.3.1

Based on its location in the periodic table, do you expect indium to be a nonmetal, a metal, or a metalloid?

Answer

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure 1.4.3.3. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

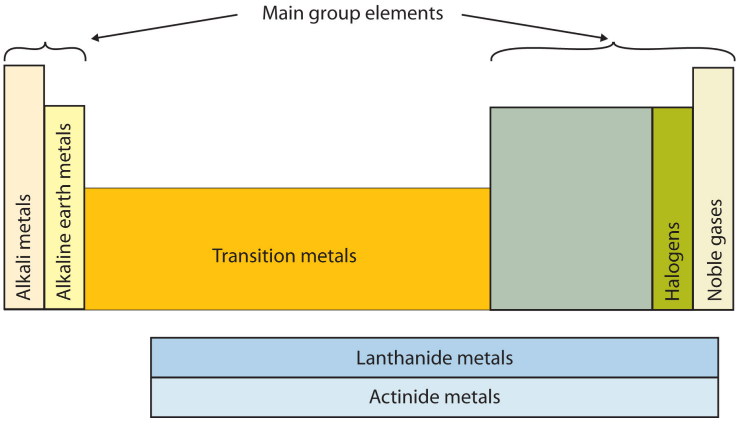


Figure 1.4.3.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth’s crust, respectively.

Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

Group 2: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for “salt forming,” which reflects that all of the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth’s ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

✓ Example 1.4.3.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

a. Lithium is an alkali metal (Group 1)
b. Argon is a noble gas (Group 18)
c. Chlorine is a halogen (Group 17)

? Exercise 1.4.3.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a:

Fluorine is a halogen (Group 17).

Answer b:

Calcium is a alkaline earth metal (Group 2).

Answer c:

Krypton is a noble gas (Group 18).

✓ Example 1.4.3.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. Li
b. Ar
c. Am
d. Fe

Solution

a. Lithium is a metal.
b. Argon is a non metal.
c. Americium is an inner transition metal.
d. Iron is a transition metal.

? Exercise 1.4.3.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. F
b. U
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

Uranium is a metal (and a inner transition metal too).

Answer c:

Copper is a metal (and a transition metal too).

Summary

The periodic table is an arrangement of the elements in order of increasing atomic number. Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods. Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.

The elements can be broadly divided into metals, nonmetals, and semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
2. Sisler, Harry H. Electronic structure, properties, and the periodic law. New york; Reinhold publishing corporation, 1963.
3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

Contributions & Attributions

1.4.4: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Atoms tend to lose, gain, or share some valance electrons, making bonds to acquire the electron configuration of the nearest noble gas, i.e., usually eight valence electrons for the main group elements, called an octet. Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while others lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called **cations**. Most metals become cations when they make ionic compounds.

Cations

A neutral sodium atom is likely to achieve an octet in its outermost shell by losing its one valence electron.

$$\mathrm{Na} \rightarrow \mathrm{Na}^{+} + \mathrm{e}^{-}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure $1.4.4.1$ is a graphical depiction of this process.

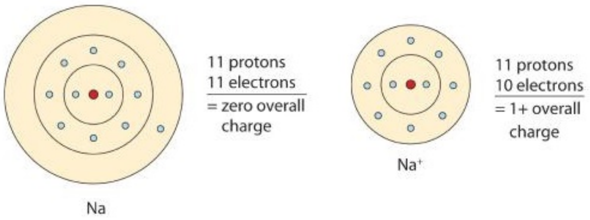


Figure $1.4.4.1$: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called **anions**. Most nonmetals become anions when they make ionic compounds.

A neutral chlorine atom has seven electrons in its outermost shell. Only one more electron is needed to achieve an octet in chlorine's valence shell. (In table salt, this electron comes from the sodium atom.)

$$\mathrm{e}^{-} + \mathrm{Cl} \longrightarrow \mathrm{Cl}^{-}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (-*ide* instead of -*ine*) to create the name of this anion. Figure $1.4.4.2$is a graphical depiction of this process.

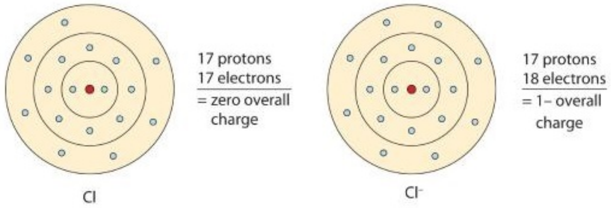


Figure 1.4.4.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall charge.

> *The names for positive and negative ions are pronounced CAT-eye-ons and ANN-eye-ons, respectively.*

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge. Figure 1.4.4.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written Ba^{2+}, not Ba^{+2}.

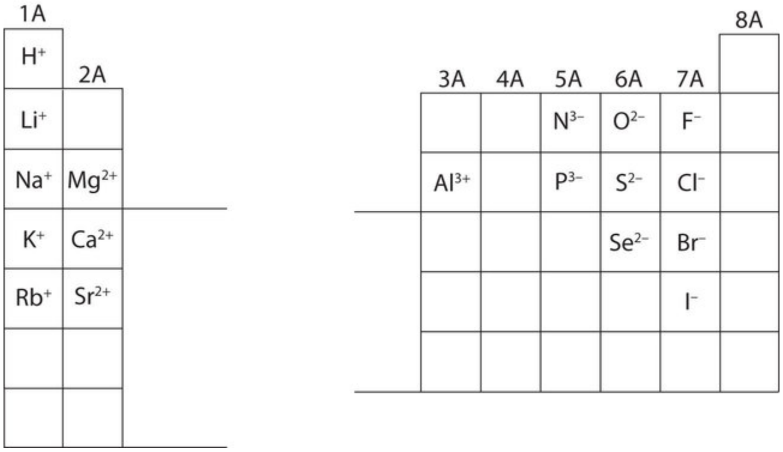


Figure 1.4.4.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

1.4.4.1: Practice Ions

Exercise 1.4.4.1.1

For each element listed, tell whether it forms a cation or an anion. Then tell whether the charge is predictable or variable. If it is predictable, give the charge.

a) O

b) Li

c) Cl

d) Ca

e) Ti

f) Se

g) Fe

h) P

i) K

j) Ag

Answer a

anion, predictable, –2

Answer b

cation, predictable, +1

Answer c

anion, predictable, –1

Answer d

cation, predictable, +2

Answer e

cation, variable

Answer f

anion, predictable, –2

Answer g

cation, variable

Answer h

anion, predictable, –3

Answer i

cation, predictable, +1

Answer j

cation, predictable, +1

Exercise 1.4.4.1.1

Which elements in the list below do not usually form ions at all?

Na Ne N C Ca Br B Si S Al

Answer

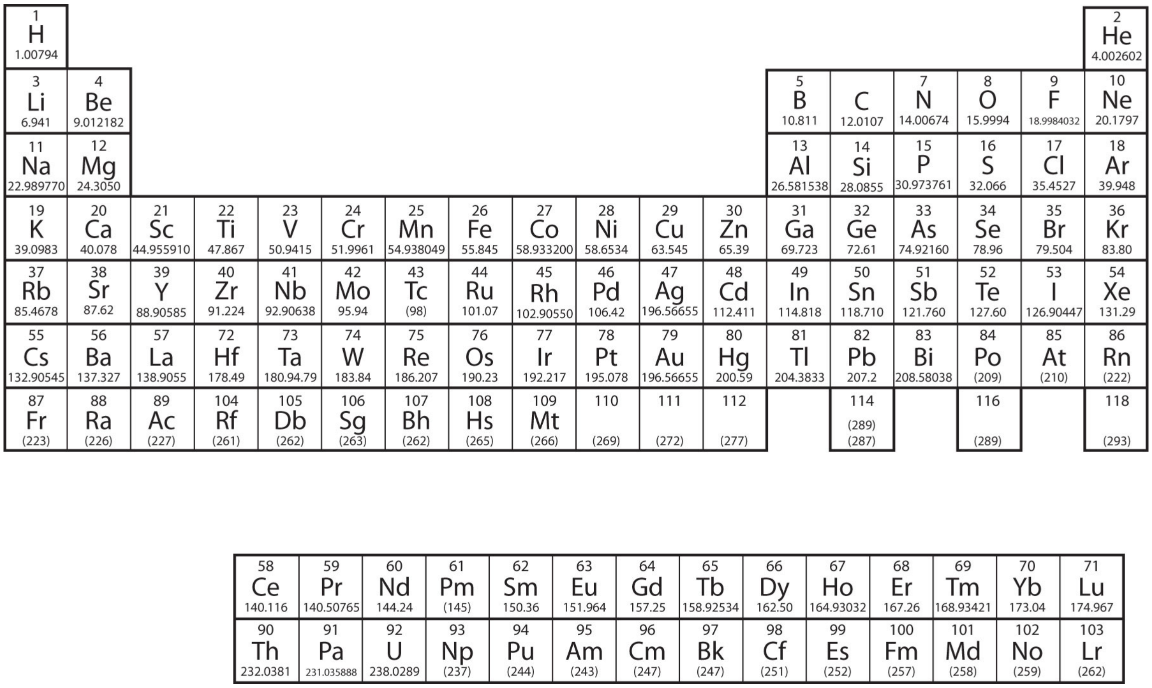
Ne C B Si

1.4.5: Electronic Structure and the Periodic Table

Learning Objectives

- Relate the electron configurations of the elements to the shape of the periodic table.
- Determine the expected electron configuration of an element by its place on the periodic table.

In Chapter 3, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure 1.4.5.1. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.



1 H 1.00794																	2 He 4.002602
3 Li 6.941	4 Be 9.012182											5 B 10.811	C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797
11 Na 22.989770	12 Mg 24.3050											13 Al 26.581538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 196.56655	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29
55 Cs 132.90545	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.94.79	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.56655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.58038	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)		114 (289) (287)		116 (289)		118 (293)

58 Ce 140.116	59 Pr 140.50765	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.92534	66 Dy 162.50	67 Ho 164.93032	68 Er 167.26	69 Tm 168.93421	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa 231.035888	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Figure 1.4.5.1: The Periodic Table

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations, the shape of the periodic table mimics the filling of the subshells with electrons.

> *The shape of the periodic table mimics the filling of the subshells with electrons.*

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure 1.4.5.2)

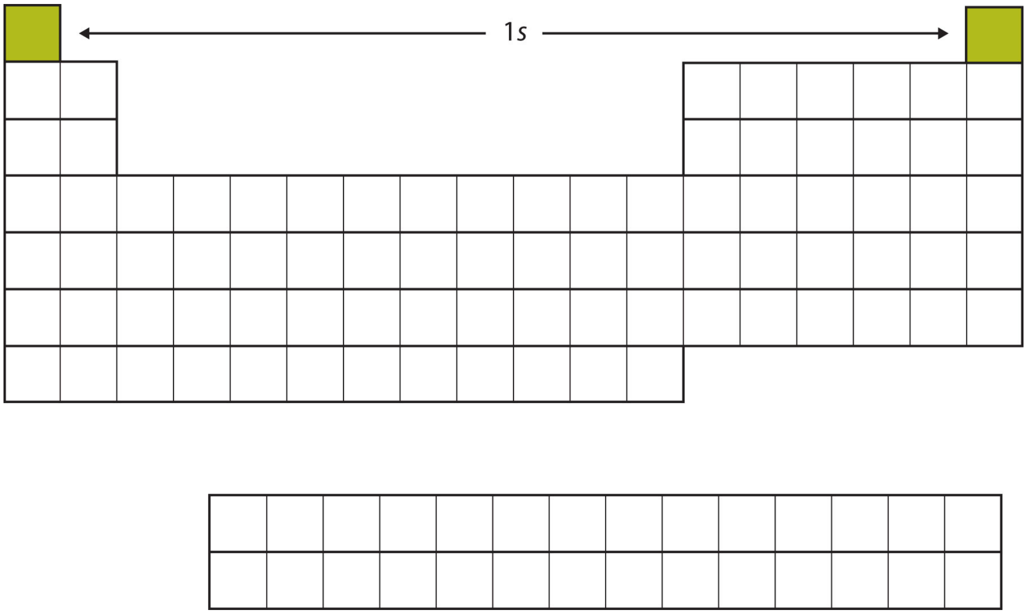


Figure 1.4.5.2: The 1s Subshell. H and He represent the filling of the 1s subshell.

The next two electrons, for Li and Be, would go into the 2*s* subshell. Figure 1.4.5.3 shows that these two elements are adjacent on the periodic table.

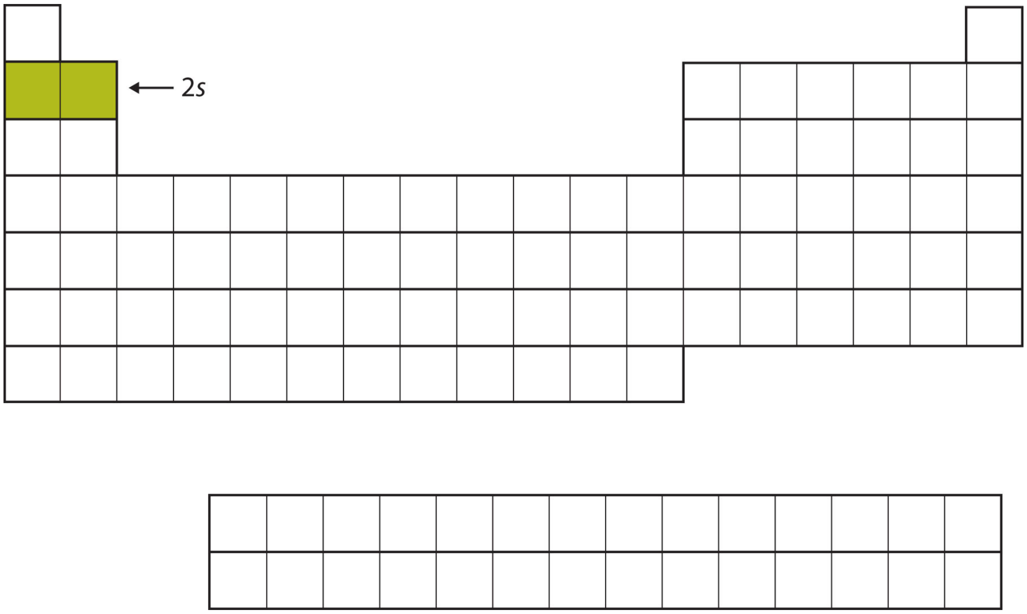


Figure 1.4.5.3: The 2s Subshell. In Li and Be, the 2s subshell is being filled.

For the next six elements, the 2*p* subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 1.4.5.4).

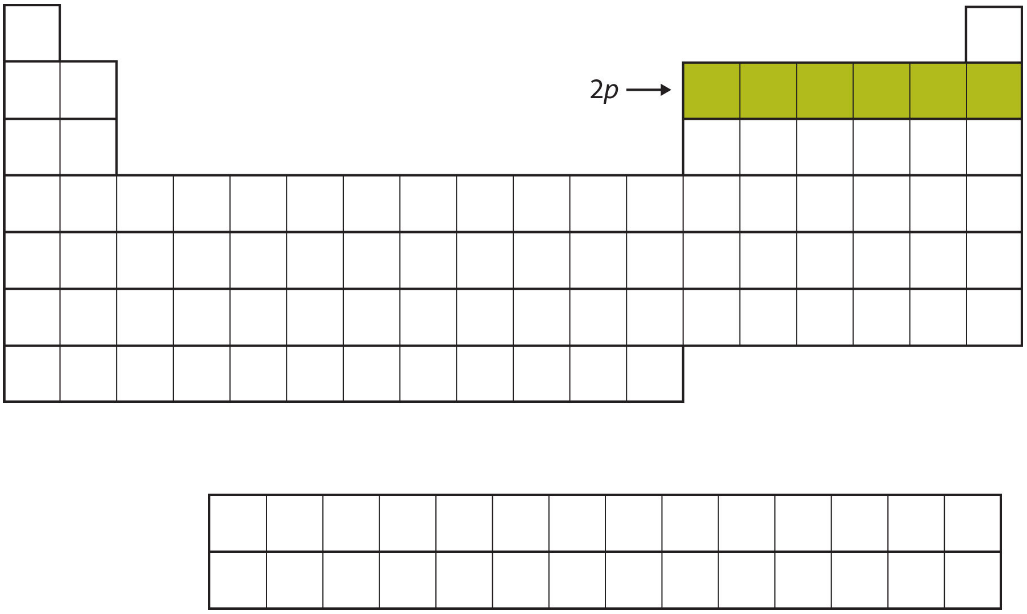


Figure 1.4.5.4: The 2p Subshell. For B through Ne, the 2p subshell is being occupied.

The next subshell to be filled is the 3*s* subshell. The elements coinciding this subshell being filled, Na and Mg, are back on the left side of the periodic table (Figure 1.4.5.5).

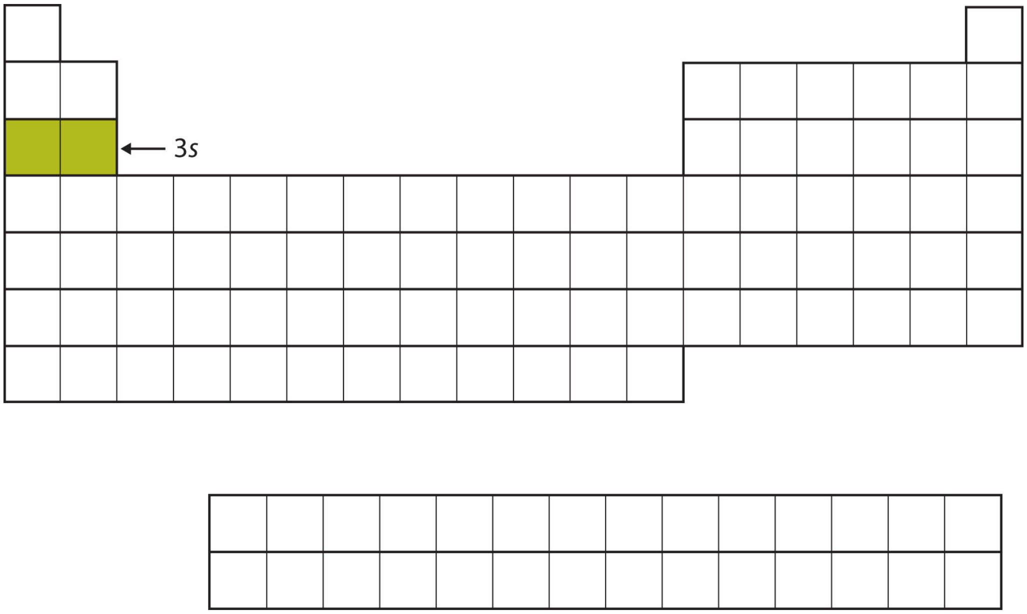


Figure 1.4.5.5: The 3s Subshell. Now the 3s subshell is being occupied.

Next, the 3*p* subshell is filled with the next six elements (Figure 1.4.5.6).

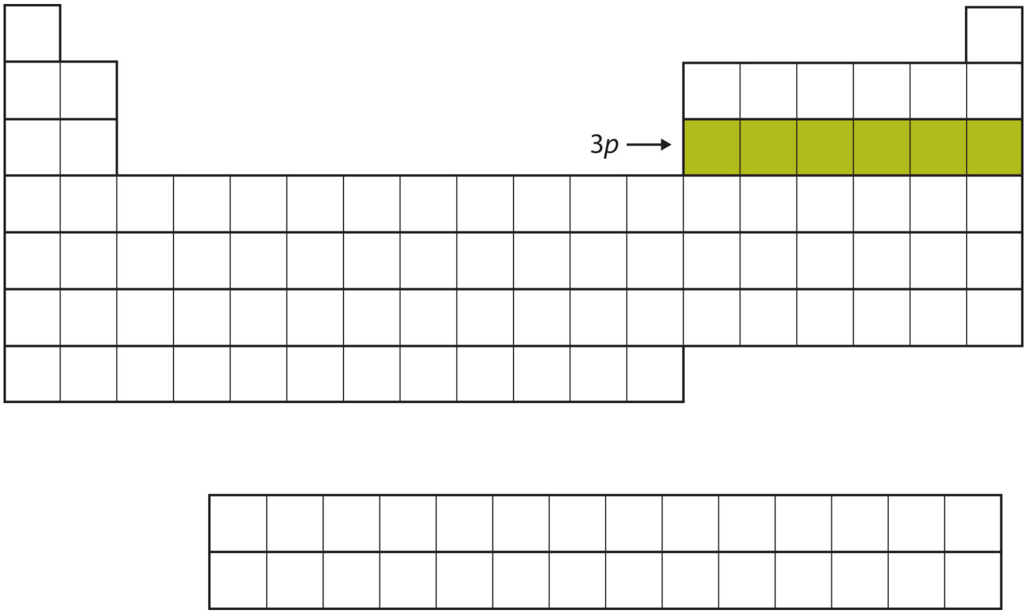


Figure8.5.6: The 3p Subshell. Next, the 3p subshell is filled with electrons.

Instead of filling the 3*d* subshell next, electrons go into the 4*s* subshell (Figure 1.4.5.7).

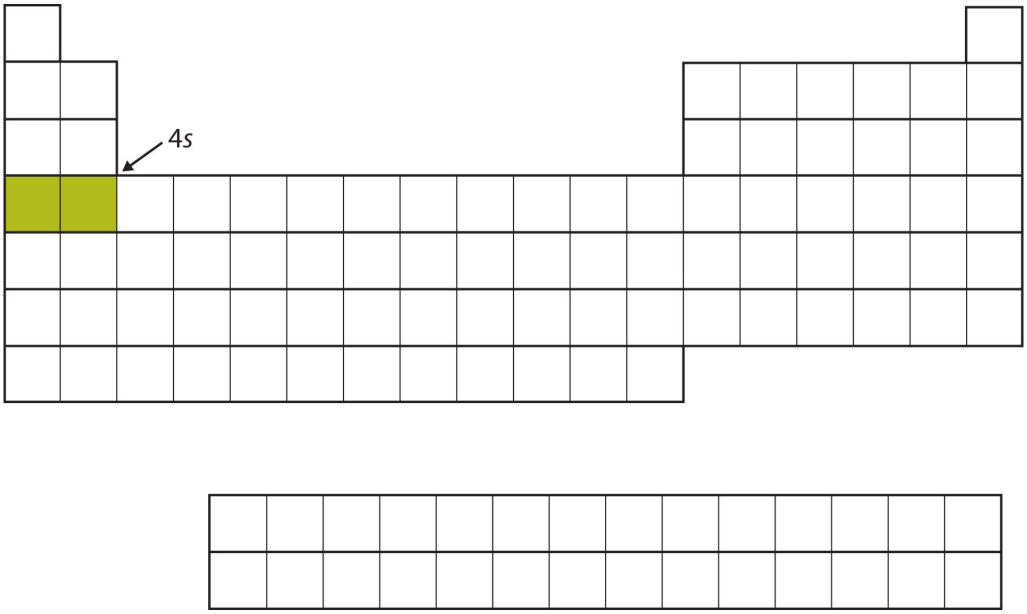


Figure 1.4.5.7: The 4s Subshell. The 4s subshell is filled before the 3d subshell. This is reflected in the structure of the periodic table.

After the 4*s* subshell is filled, the 3*d* subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure 1.4.5.8).

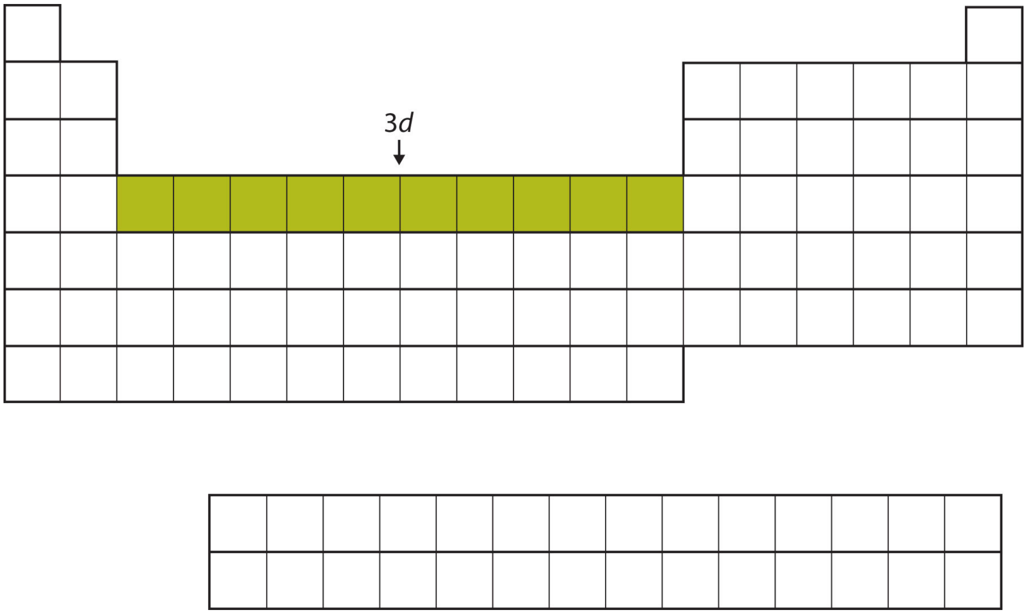


Figure 1.4.5.8: The 3d Subshell. The 3d subshell is filled in the middle section of the periodic table.

And so forth. As we go across the columns of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the *s* subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** is the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 1.4.5.9shows the blocks of the periodic table.

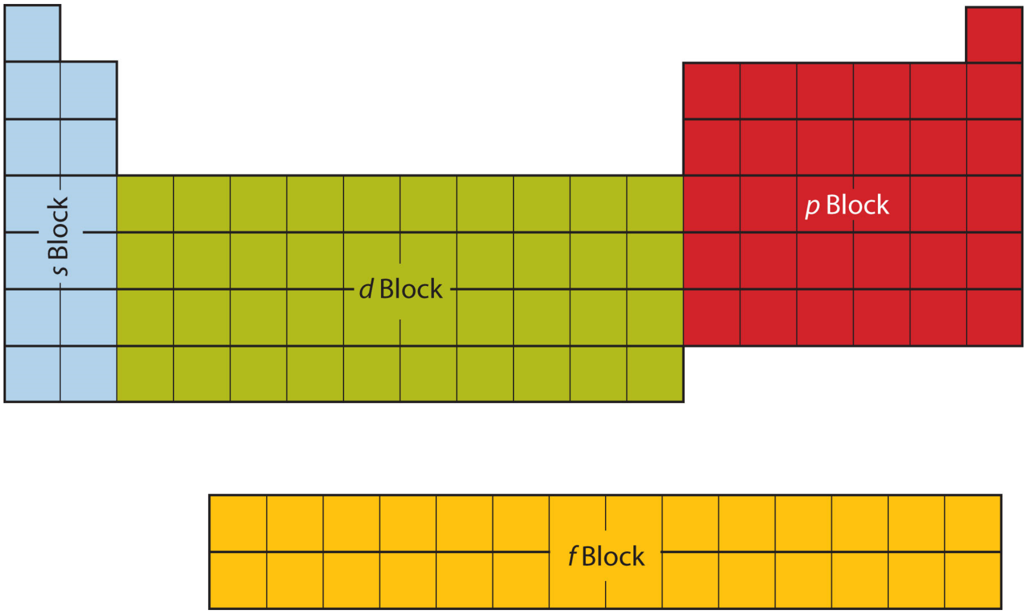


Figure 1.4.5.9: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

Empty periodic table, separated into 4 different colored blocks. S block is blue, d block is green, p block is red and f block is orange.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

Table with two columns and 6 rows. The first column on the left has various elements in the rows underneath. The second column on the right has the different corresponding electron configurations for the specified element in the rows underneath.

H:	$1s^{\mathbf{1}}$
Li:	$1s^22s^{\mathbf{1}}$
Na:	$[Ne]3s^{\mathbf{1}}$
K:	$[Ar]4s^{\mathbf{1}}$
Rb:	$[Kr]5s^{\mathbf{1}}$
Cs:	$[Xe]6s^{\mathbf{1}}$

They all have a similar electron configuration in their valence shells: a single *s* electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry—*and they do*. The organization of electrons in atoms explains not only the shape of the periodic table, but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the *s* and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure 1.4.5.10. It is in the fourth column of the *p* block. This means that its electron configuration should end in a p^4 electron configuration. Indeed, the electron configuration of Se is $[Ar]4s^23d^{10}4p^4$, as expected.

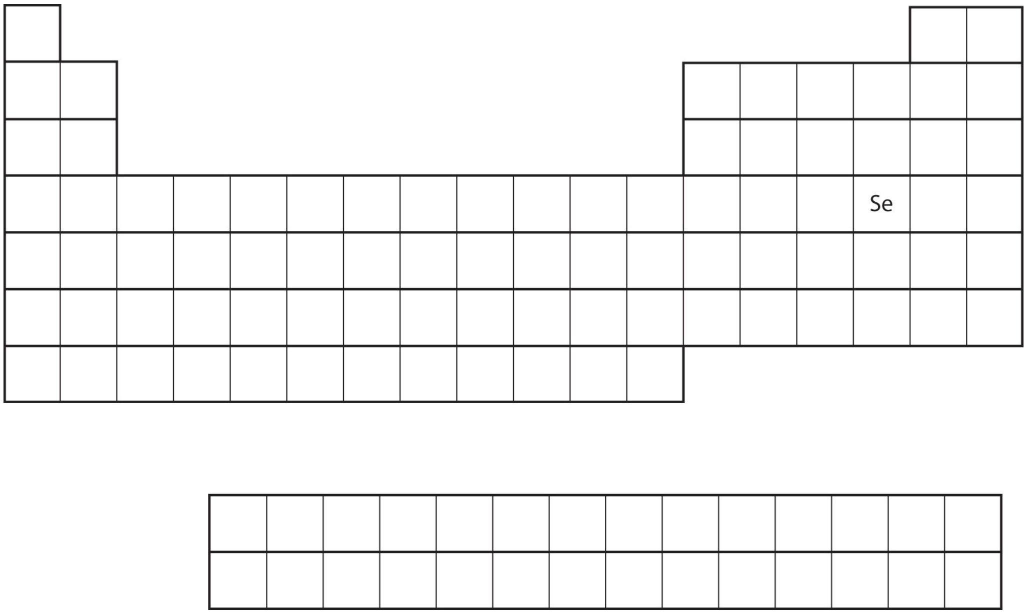


Figure 1.4.5.10: Selenium on the Periodic Table

Example 1.4.5.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure 1.4.5.11).

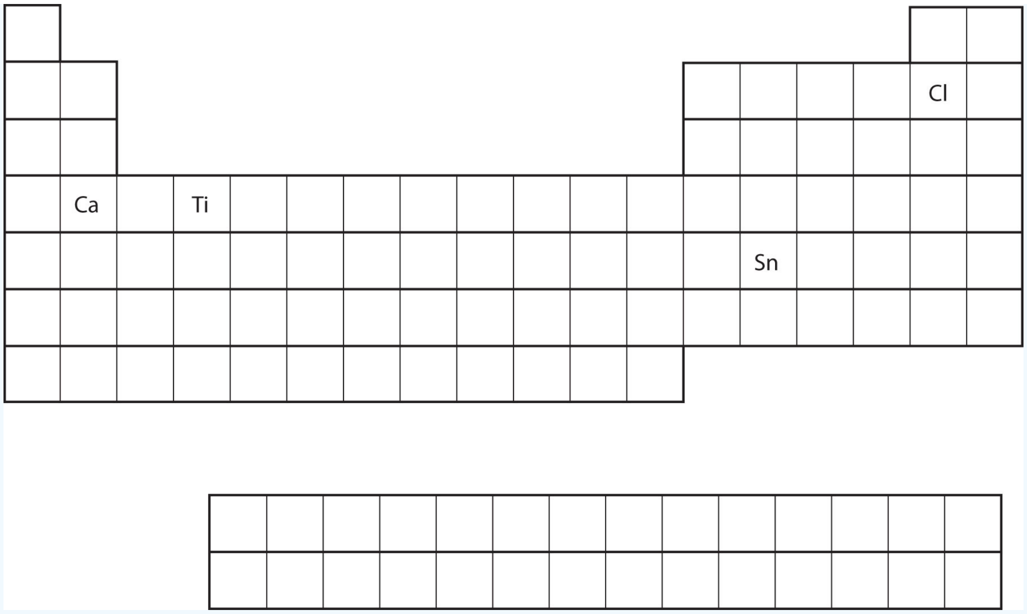


Figure 1.4.5.11: Various Elements on the Periodic Table

1. Ca
2. Sn

Solution

1. Ca is located in the second column of the *s* block. We would expect that its electron configuration should end with s^2. Calcium's electron configuration is $[Ar]4s^2$.
2. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in p^2. Tin's electron configuration is $[Kr]5s^2 4d^{10} 5p^2$.

Exercise 1.4.5.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom. Figure 1.4.5.11

a. Ti
b. Cl

Answer a

$[Ar]4s^2 3d^2$

Answer b

$[Ne]3s^2 3p^5$

Food and Drink Application: Artificial Colors

The color of objects comes from a different mechanism than the colors of neon and other discharge lights. Although colored lights produce their colors, objects are colored because they preferentially reflect a certain color from the white light that shines on them. A red tomato, for example, is bright red because it reflects red light while absorbing all the other colors of the rainbow.

Many foods, such as tomatoes, are highly colored; in fact, the common statement "you eat with your eyes first" is an implicit recognition that the visual appeal of food is just as important as its taste. But what about processed foods?

Many processed foods have food colorings added to them. There are two types of food colorings: natural and artificial. Natural food colorings include caramelized sugar for brown; annatto, turmeric, and saffron for various shades of orange or yellow;

betanin from beets for purple; and even carmine, a deep red dye that is extracted from the cochineal, a small insect that is a parasite on cacti in Central and South America. (That's right—you may be eating bug juice!)

Some colorings are artificial. In the United States, the Food and Drug Administration currently approves only seven compounds as artificial colorings in food, beverages, and cosmetics:

- FD&C Blue #1: Brilliant Blue FCF
- FD&C Blue #2: Indigotine
- FD&C Green #3: Fast Green FCF
- RD&C Red #3: Erythrosine
- FD&C Red #40: Allura Red AC
- FD&C Yellow #5: Tartrazine
- FD&C Yellow #6: Sunset Yellow FCF

Lower-numbered colors are no longer on the market or have been removed for various reasons. Typically, these artificial colorings are large molecules that absorb certain colors of light very strongly, making them useful even at very low concentrations in foods and cosmetics. Even at such low amounts, some critics claim that a small portion of the population (especially children) is sensitive to artificial colorings and urge that their use be curtailed or halted. However, formal studies of artificial colorings and their effects on behavior have been inconclusive or contradictory. Many people continue to enjoy processed foods with artificial coloring (like those shown in the accompanying figure).

Artificial food colorings. Artificial food colorings are found in a variety of food products, such as processed foods, candies, and egg dyes. Even pet foods have artificial food coloring in them. (CC by 2.0, Matthew Bland via Flickr.)

Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.

1.4.5.1: Practice Electronic Structure

Exercise 1.4.5.1.1

Which orbital is the next to fill after the 4s orbital?

Answer

3d orbital

Exercise 1.4.5.1.1

Which orbital sublevel does NOT exist?

a) 3d

b) 4f

c) 2d

d) all exist

Answer

c (The main level 2 only has 2s and 2p sublevels.)

Exercise 1.4.5.1.1

How many electrons can fit in the 4d sublevel?

Answer

10 electrons

Exercise 1.4.5.1.1

Give the electron configuration for each of the following atoms.

a) Zn

b) K

c) Br

Answer a

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$ (total of 30 electrons)

Answer b

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$ (total of 19 electrons)

Answer c

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$ (total of 35 electrons)

Exercise 1.4.5.1.1

Give the electron configuration for each of the following ions.

d) O^{2-}

e) K^{1+}

f) Br^{1-}

Answer d

$1s^2\ 2s^2\ 2p^6$ (total of 10 electrons)

Answer e

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ (total of 18 electrons)

Answer f

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$ (total of 36 electrons)

1.4.6: Periodic Trends- Atomic Radius

Figure 1.4.6.1 (Credit: James Cridland; Source: http://www.flickr.com/photos/jamescridland/613445810/(opens in new window); License: CC by 2.0(opens in new window))

How can all of these people fit in such a small space?

Events draw large numbers of people to them. Even an outdoor event can fill up so that there is no room for more people. The crowd capacity depends on the amount of space in the venue, and the amount of space depends on the size of the objects filling it. We can get more people into a given space than can elephants, because elephants are larger than people. We can get more squirrels into that same space than we can people for the same reason. Knowing the sizes of objects to be dealt with can be important in deciding how much space is needed.

Atomic Radius

The size of atoms is important to explanations of the behavior of atoms or compounds. One way to express the size of atoms is by use of **atomic radius**. This data helps us understand why some molecules fit together and why other molecules have parts that get too crowded under certain conditions.

The size of an atom is defined by the edge of its orbital. However, orbital boundaries are fuzzy, and variable under different conditions. In order to standardize the measurement of atomic radii, the distance between the nuclei of two identical atoms bonded together is measured. The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together.

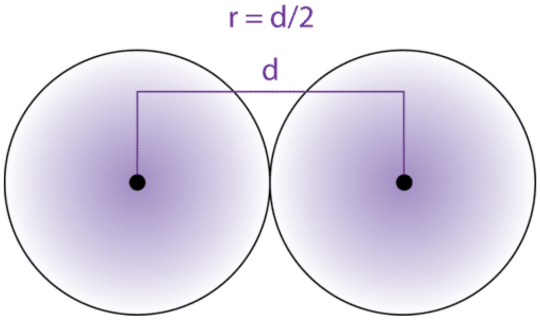


Figure 1.4.6.2: The atomic radius (r) of an atom can be defined as one half the distance (d) between two nuclei in a diatomic molecule. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Atomic radii have been measured for elements. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be 74 pm. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2}=37\text{ pm}$.

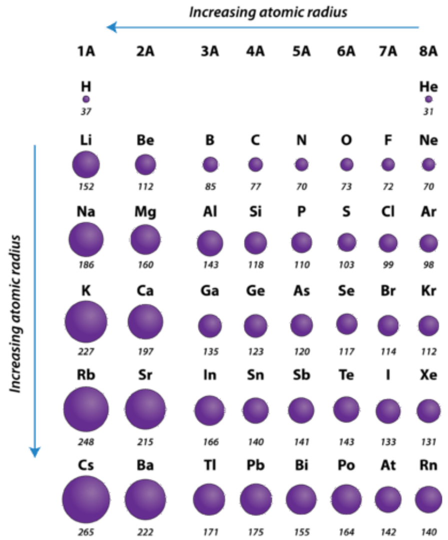


Figure 1.4.6.3: Atomic radii of the representative elements measured in picometers. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Periodic Trend

The atomic radius of atoms generally decreases from left to right across a period. There are some small exceptions, such as the oxygen radius being slightly greater than the nitrogen radius. Within a period, protons are added to the nucleus as electrons are being added to the same principal energy level. These electrons are gradually pulled closer to the nucleus because of its increased positive charge. Since the force of attraction between nuclei and electrons increases, the size of the atoms decreases. The effect lessens as one moves further to the right in a period, because of electron-electron repulsions that would otherwise cause the atom's size to increase.

Group Trend

The atomic radius of atoms generally increases from top to bottom within a group. As the atomic number increases down a group, there is again an increase in the positive nuclear charge. However, there is also an increase in the number of occupied principal energy levels. Higher principal energy levels consist of orbitals which are larger in size than the orbitals from lower energy levels. The effect of the greater number of principal energy levels outweighs the increase in nuclear charge, and so atomic radius increases down a group.

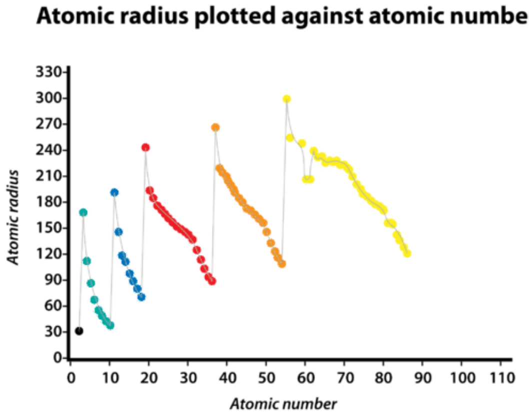


Figure 1.4.6.4: A graph of atomic radius plotted versus atomic number. Each successive period is shown in a different color. As the atomic number increases within a period, the atomic radius decreases. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Summary

- Atomic radius is determined as half the distance between the nuclei of two identical atoms bonded together.
- The atomic radius of atoms generally decreases from left to right across a period.
- The atomic radius of atoms generally increases from top to bottom within a group.

Review

1. Define “atomic radius.”
2. What are the units of measurement for atomic radius?
3. How does the atomic radius of different elements change across a period?
4. How does atomic radius change from top to bottom within a group?
5. Explain why the atomic radius of hydrogen is so much smaller than the atomic radius of potassium.

- **6.15: Periodic Trends- Atomic Radius** by CK-12 Foundation is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

1.4.6.1: Periodic Trends- Ionization Energy

Figure 1.4.6.1.1 (Credit: Nick Russill; Source: Flickr, Northern Lights, Greenland(opens in new window) [www.flickr.com]; License: CC by 2.0(opens in new window))

The incredible green lights in this cold northern sky consist of charged particles known as ions. Their swirling pattern is caused by the pull of Earth’s magnetic field. Called the northern lights, this phenomenon of nature shows that ions respond to a magnetic field. Do you know what ions are? Read on to find out.

Atoms Are Neutral

The northern lights aren’t caused by atoms, because atoms are not charged particles. An atom always has the same number of electrons as protons. Electrons have an electric charge of -1 and protons have an electric charge of +1. Therefore, the charges of an atom’s electrons and protons “cancel out.” This explains why atoms are neutral in electric charge.

Q: What would happen to an atom’s charge if it were to gain extra electrons?

A: If an atom were to gain extra electrons, it would have more electrons than protons. This would give it a negative charge, so it would no longer be neutral.

Atoms to Ions

Atoms cannot only gain extra electrons. They can also lose electrons. In either case, they become **ions**. Ions are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons. If atoms lose electrons, they become positive ions, or cations. If atoms gain electrons, they become negative ions, or anions. Consider the example of fluorine (see figure below). A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with an electric charge of -1.

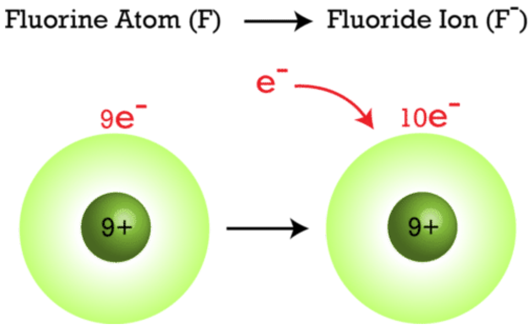


Figure 1.4.6.1.2 (Credit: Christopher AuYeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Names and Symbols

Like fluoride, other negative ions usually have names ending in *–ide*. Positive ions, on the other hand, are just given the element name followed by the word *ion*. For example, when a sodium atom loses an electron, it becomes a positive sodium ion. The charge of an ion is indicated by a plus (+) or minus sign (-), which is written to the right of and just above the ion’s chemical symbol. For example, the fluoride ion is represented by the symbol F^-, and the sodium ion is represented by the symbol Na^+. If the charge is greater than one, a number is used to indicate it. For example, iron (Fe) may lose two electrons to form an ion with a charge of plus two. This ion would be represented by the symbol Fe^{2+}. This and some other common ions are listed with their symbols in the table below.

Table $PageIndex1$: Some Common Ions

Cations		Anions	
Name of Ion	Chemical Symbol	Name of Ion	Chemical Symbol
Calcium ion	Ca^{2+}	Chloride	Cl^{-}
Hydrogen ion	H^{+}	Fluoride	F^{-}
Iron(II) ion	Fe^{2+}	Bromide	Br^{-}
Iron(III) ion	Fe^{3+}	Oxide	O^{2-}

Q: How does the iron(III) ion differ from the iron(II) ion?

A: The iron(III) ion has a charge of +3, so it has one less electron than the iron(II) ion, which has a charge of +2.

Q: What is the charge of an oxide ion? How does its number of electrons compare to its number of protons?

A: An oxide ion has a charge of -2. It has two more electrons than protons.

How Ions Form

The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation. The radiation may give their outer electrons enough energy to escape from the attraction of the positive nucleus. However, most ions form when atoms transfer electrons to or from other atoms or molecules. For example, sodium atoms may transfer electrons to chlorine atoms. This forms positive sodium ions (Na^{+}) and negative chloride ions (Cl^{-}).

Q: Why do you think atoms lose electrons to, or gain electrons from, other atoms?

A: Atoms form ions by losing or gaining electrons because it makes them more stable and this state takes less energy to maintain. The most stable state for an atom is to have its outermost energy level filled with the maximum possible number of electrons. In the case of metals such as lithium, with just one electron in the outermost energy level, a more stable state can be achieved by losing that one outer electron. In the case of nonmetals such as fluorine, which has seven electrons in the outermost energy level, a more stable state can be achieved by gaining one electron and filling up the outer energy level.

Properties of Ions

Ions are highly reactive, especially as gases. They usually react with ions of opposite charge to form neutral compounds. For example, positive sodium ions and negative chloride ions react to form the neutral compound sodium chloride, commonly known as table salt. This occurs because oppositely charged ions attract each other. Ions with the same charge, on the other hand, repel each other. Ions are also deflected by a magnetic field, as you saw in the opening image of the northern lights.

Summary

- Atoms have equal numbers of positive protons and negative electrons, so they are neutral in electric charge.
- Atoms can gain or lose electrons and become ions, which are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons.
- The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation or when atoms transfer electrons to or from other atoms.
- Ions are reactive, attracted or repulsed by other charged particles, and deflected by a magnetic field.

Review

1. Why are atoms neutral in electric charge?
2. Define ion.
3. Compare and contrast cations and anions, and give an example of each.
4. Describe how ions form.
5. List properties of ions.
6. The model in the illustration below represents an atom of lithium (Li). If the lithium atom becomes an ion, which type of ion will it be, a cation or an anion? What will be the electric charge of this ion? What will the ion be named? What symbol will be used to represent it?

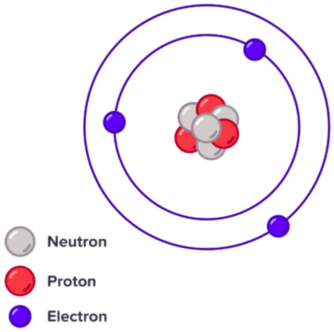


Figure 1.4.6.1.3 (Credit: CK-12 Foundation Employee; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

1.4.6.2: Metallic and Nonmetallic Character

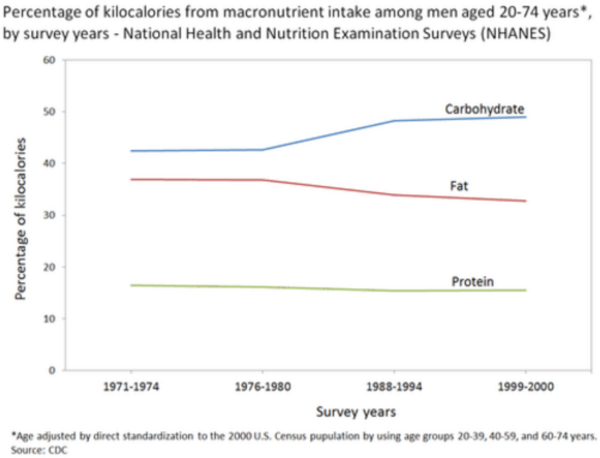


Figure 1.4.6.2.1: Graph indicating eating habits of American men from 1971 to 2000. (Credit: Recreated by CK-12 Foundation based on data from the CDC; Source: CDC GOV Preview(opens in new window) [www.cdc.gov]; Flickr, Venn Diagram(opens in new window) [www.flickr.com]; License: Public Domain)

What are we eating as a nation?

The graph above indicates some trends in the U.S. diet over a thirty-year period. By observing the direction our eating habits are going, steps can be taken to help prevent bad eating habits and decrease problems such as high blood pressure and heart attacks.

Development of the periodic table has helped organize chemical information in many ways. We can now see trends among properties of different atoms and make predictions about the behavior of specific materials.

Metallic and Nonmetallic Character

Metallic character refers to the level of reactivity of a metal. Metals tend to lose electrons in chemical reactions, as indicated by their low ionization energies. Within a compound, metal atoms have relatively low attraction for electrons, as indicated by their low electronegativities. By following the trend summary in the figure below, you can see that the most reactive metals would reside in the lower left portion of the periodic table. The most reactive metal is cesium, which is not found in nature as a free element. It reacts explosively with water and will ignite spontaneously in air. Francium is below cesium in the alkali metal group, but is so rare that most of its properties have never been observed.

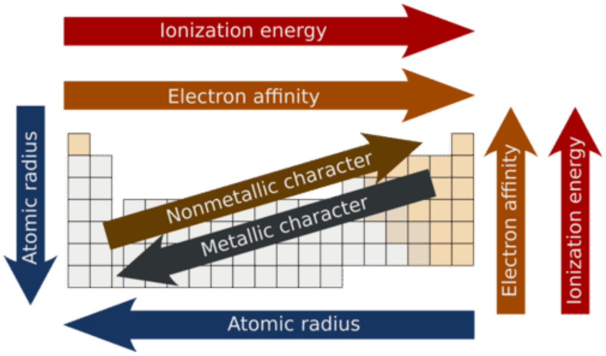


Figure 1.4.6.2.2: Trends in behaviors of elements. (Credit: User:Mirek2/Wikimedia Commons; Source: Commons Wikimedia, Periodic Trends(opens in new window) [commons.wikimedia.org]; License: Public Domain)

Reactivity of metals is based on processes such as the formation of halide compounds with halogens, and how easily the element displaces hydrogen from dilute acids.

The metallic character increases as you go down a group. Since the ionization energy decreases going down a group (or increases going up a group), the increased ability for metals lower in a group to lose electrons makes them more reactive. In addition, the atomic radius increases going down a group, placing the outer electrons further away from the nucleus and making that electron less attracted by the nucleus.

Nonmetals tend to gain electrons in chemical reactions, and have a high attraction for electrons within a compound. The most reactive nonmetals reside in the upper right portion of the periodic table. Since the noble gases are a special group because of their lack of reactivity, the element fluorine is the most reactive nonmetal. It is not found in nature as a free element. Fluorine gas reacts explosively with many other elements and compounds, and is considered to be one of the most dangerous known substances.

Note that there is no clear division between metallic and **non-metallic** character. As we move across the periodic table, there is an increasing tendency to accept electrons (nonmetallic) and a decrease in the possibility that an atom will give up one or more electrons.

Summary

- Metallic character refers to the level of reactivity of a metal.
- Non-metallic character relates to the tendency to accept electrons during chemical reactions.
- Metallic tendency increases going down a group.
- Non-metallic tendency increases going from left to right across the periodic table.

Review

1. Define “metallic character.”
2. Define “non-metallic character.”
3. Describe the trend in metallic character going down a group.
4. Describe the trend in non-metallic character going across the periodic table.
5. Why does the metallic character increase as you go down a group?

- **6.22: Periodic Trends - Metallic and Nonmetallic Character** by CK-12 Foundation is licensed CK-12. Original source: https://flexbooks.ck12.org/cbook/ck-12-chemistry-flexbook-2.0/.

1.4.6.3: Practice Trends

Exercise 1.4.6.3.1

Choose the element in each pair that has a larger atomic radius (size).

a) S or Ar

b) Na or K

c) O or P

Answer a

S

Answer b

K

Answer c

P

Exercise 1.4.6.3.1

Choose the element in each pair that has a greater ionization energy.

d) Sc or K

e) Ca or Mg

f) Al or P

Answer d

Sc

Answer e

Mg

Answer f

P

Exercise 1.4.6.3.1

Choose the element in each pair that has more metallic character.

g) Si or Al

h) Ca or Ba

i) Ti or Fe

Answer g

Al

Answer h

Ba

Answer i

Ti

Exercise 1.4.6.3.1

C F

Ge Br

Out of the four elements listed above, choose the one that has

j) the smallest radius

k) the least ionization energy

l) the most metallic character

Answer j

F

Answer k

Ge

Answer l

Ge

1.4.7: Calculating Atomic Mass

Figure 1.4.7.1 (Credit: George M. Groutas; Source: Flickr Boulders beach, Cape Peninsula(opens in new window) [www.flickr.com]; License: CC by 2.0(opens in new window))

Have you ever tried to move a boulder?

You have a pile of rocks to move and need to decide what equipment you want to rent to move them. If the rocks are fairly small, you can get a shovel to pick them up. Larger rocks could be moved by hand, but big boulders will need some sort of mechanical scoop. The amount of each kind of rock will also determine how much time you will need to get the job done. Knowing the relative amounts of large, medium, and small rocks can be very useful in deciding how to approach the job.

Percent Natural Abundance

Most elements occur naturally as a mixture of two or more isotopes. The table below shows the natural isotopes of several elements, along with the **percent natural abundance** of each.

Table 1.4.7.1: Atomic Masses and Percents of Abundance of Some Natural Isotopes

Element	**Isotope (Symbol)**	**Percent Natural Abundance**	**Atomic Mass** (amu)	**Average Atomic Mass** (amu)
Hydrogen	${}^{1}_{1}\text{H}$	99.985	1.0078	1.0080
	${}^{2}_{1}\text{H}$	0.015	2.0141	
	${}^{3}_{1}\text{H}$	negligible	3.0160	
Carbon	${}^{12}_{6}\text{C}$	98.89	12.000	12.011
	${}^{13}_{6}\text{C}$	1.11	13.003	
	${}^{14}_{6}\text{C}$	trace	14.003	
Oxygen	${}^{16}_{8}\text{O}$	99.759	15.995	15.999
	${}^{17}_{8}\text{O}$	0.037	16.995	
	${}^{18}_{8}\text{O}$	0.204	17.999	
Chlorine	${}^{35}_{17}\text{Cl}$	75.77	34.969	35.453
	${}^{37}_{17}\text{Cl}$	24.23	36.966	
Copper	${}^{63}_{29}\text{Cu}$	69.17	62.930	63.546
	${}^{65}_{29}\text{Cu}$	30.83	64.928	

For some elements, one particular isotope predominates greatly over the other isotopes. Naturally occurring hydrogen is nearly all hydrogen-1 and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in more substantial quantities. Chlorine (atomic number 17) is a yellowish-green toxic gas. About three quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise **atomic masses**, you would get 36.

$$\frac{(34.969 + 36.966)}{2} = 35.968 \text{ amu}$$

Clearly the actual average atomic mass from the last column of the table is significantly lower. Why? We need to take into account the percent natural abundance of each isotope, in order to calculate the weighted average. The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element. The sample problem below demonstrates how to calculate the atomic mass of chlorine.

Example 1.4.7.1

Use the atomic masses of each of the two isotopes of chlorine along with their respective percent abundances to calculate the average atomic mass of chlorine.

Solution

Step 1: List the known and unknown quantities and plan the problem.

Known

- Chlorine-35: atomic mass $= 34.969 \text{ amu}$and percent abundance $= 75.77\%$
- Chlorine-37: atomic mass $= 36.966 \text{ amu}$and percent abundance $= 24.23\%$

Unknown

- Average atomic mass of chlorine

Change each percent abundance into decimal form by dividing by 100. Multiply this value by the atomic mass of that isotope. Add together for each isotope to get the average atomic mass.

Step 2: Calculate.

$$\begin{array}{ll} \text{chlorine-35} & 0.7577 \times 34.969 = 26.50 \text{ amu} \\ \text{chlorine-37} & 0.2423 \times 36.966 = 8.957 \text{ amu} \\ \text{average atomic mass} & 26.50 + 8.957 = 35.46 \text{ amu} \end{array}$$

Note: Applying significant figure rules results in the 35.45 amuresult without excessive rounding error. In one step:

$$(0.7577 \times 34.969) + (0.2423 \times 36.966) = 35.46 \text{ amu}$$

Step 3: Think about your result.

The calculated average atomic mass is closer to 35 than to 37 because a greater percentage of naturally occurring chlorine atoms have the mass number of 35. It agrees with the value from the table above.

Summary

- The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element.
- Calculations of atomic mass use the percent abundance of each isotope.

Review

1. Define atomic mass.
2. What information do you need to calculate atomic mass for an element?
3. Calculate the atomic mass for carbon using the data provided in the table below.

Isotope	Atomic Mass	Percent Abundance
carbon-12	12.000000	98.90
carbon-13	13.003355	1.100

1.4.7.1: Practice Atomic Mass

Exercise 1.4.7.1.1

The element lithium is found in nature as a mixture of ^{6}Li and ^{7}Li. Given that the average atomic mass of lithium is 6.94 amu, which isotope is more abundant?

Answer

^{7}Li is more abundant. Since the average mass is really close to 7 amu, you have a lot more of the isotope with a mass of 7 amu than you do of the isotope with a mass of 6 amu.

Exercise 1.4.7.1.1

The element strontium has the following natural abundances. Calculate the average atomic mass.

82.58 % is ^{88}Sr with mass of 87.9056 amu

9.86 % is ^{86}Sr with mass of 85.9093 amu

7.00 % is ^{87}Sr with mass of 86.9089 amu

0.56 % is ^{84}Sr with mass of 83.9134 amu

Answer

Add up (uncertain digits are underlined):

$72.5\underline{9}244$ amu

$8.4\underline{7}066$ amu

$6.0\underline{8}362$ amu

$0.4\underline{6}992$ amu

Total is $87.6\underline{1}664$ amu which rounds to 87.62 amu

Exercise 1.4.7.1.1

If there were chemists on another planet, their periodic table might be different because the elements might have different "natural" abundances. If on the other planet, the element titanium had the following natural abundances, calculate the average atomic mass that the chemists on that planet would use.

73.64 % is ^{49}Ti with mass of 48.9479 amu

18.82 % is ^{50}Ti with mass of 49.9448 amu

7.54 % is ^{48}Ti with mass of 47.9479 amu

Answer

Add up (uncertain digits are underlined):

$36.0\underline{4}523$ amu

$9.39\underline{9}61$ amu

$3.6\underline{1}527$ amu

Total is $49.0\underline{60}11$ amu which rounds to 49.06 amu

Calculating the other way:

I realized that this is not one our our goals for Chem 142. We do this in Chem 101. I left the problems here anyway in case you are curious.

Exercise 1.4.7.1.1

The element antimony is made up of two isotopes, ^{121}Sb with a mass of 120.9038 amu and ^{123}Sb with a mass of 122.9042 amu. Given that the average atomic mass is 121.760 amu, what are the percent abundances of the two isotopes?

Hint

The average atomic mass is equal to the sum of the fractional abundance times mass of each isotope.

average = (fractional abundance of ^{121}Sb)(mass of ^{121}Sb) + (fractional abundance of ^{123}Sb)(mass of ^{123}Sb)

Set the fractional abundance of ^{121}Sb = x. Since there are only 2 isotopes, and they must add up to 1 whole, the fractional abundance of ^{123}Sb = 1 – x

Answer

121.760 amu = (x)(120.9038 amu) + (1 – x)(122.9042 amu)

121.760 amu = (x)(120.9038 amu) + 122.9042 amu – (x)(122.9042 amu) now subtract 122.9042 amu from both sides

– 1.14$\underline{42}$ amu = (x)(120.9038 amu) – (x)(122.9042 amu) combine terms on right

– 1.14$\underline{42}$ amu = (x)(–2.000$\underline{4}$ amu) divide both by – 2.0004 amu

0.571$\underline{9}$86 = x

So fractional abundance of ^{121}Sb is 0.5720, and percent abundance is 57.20 %.

Fractional abundance of ^{123}Sb is therefore 0.4280, and percent abundance is 42.80 %

Exercise 1.4.7.1.1

On our imaginary planet, the element zinc is made up of two isotopes, ^{66}Zn with a mass of 65.9260 amu and ^{68}Zn with a mass of 67.9248 amu. Given that their average atomic mass for zinc is 66.840 amu, what are the percent abundances of the two isotopes on that planet?

Answer

66.840 amu = (x)(65.9620 amu) + (1 – x)(67.9248 amu)

66.840 amu = (x)(65.9620 amu) + 67.9248 amu – (x)(67.9248 amu) now subtract 67.9248 amu from both sides

– 1.08$\underline{48}$ amu = (x)(65.9620 amu) – (x)(67.9248 amu) combine terms on right

– 1.08$\underline{48}$ amu = (x)(–1.962$\underline{8}$ amu) divide both by – 1.9628 amu

0.552$\underline{6}$798 = x

So fractional abundance of ^{66}Zn is 0.5527, and percent abundance is 55.27 %.

Fractional abundance of ^{68}Zn is therefore 0.4473, and percent abundance is 44.73 %

SECTION OVERVIEW

1.5: Nuclear Chemistry

1.5.1: Radioactivity

Learning Objectives

- To define and give examples of the major types of radioactivity.

Atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of the atom, while electrons circle the nucleus in shells and subshells and account for an atom's size. Remember, the notation for succinctly representing an isotope of a particular atom:

$$^{12}_{6}\mathrm{C} \tag{1.5.1.1}$$

The element in this example, represented by the symbol C, is carbon. Its atomic number, 6, is the lower left subscript on the symbol and is the number of protons in the atom. The mass number, the superscript to the upper left of the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 12, which means that the number of neutrons in the atom is 12 − 6 = 6 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neutrons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. The two isotopes of hydrogen, $^{2}\mathrm{H}$ and $^{3}\mathrm{H}$, are given their own names: deuterium (D) and tritium (T), respectively. Another way of expressing a particular isotope is to list the mass number after the element name, like carbon-12 or hydrogen-3.

Atomic theory in the 19th century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source as the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an alpha particle, which is symbolized by the Greek letter α. An alpha particle is composed of two protons and two neutrons, and so it is the same as a helium nucleus. (We often use $^{4}_{2}\mathrm{He}$ to represent an alpha particle.) It has a 2+ charge. When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by two (because of the loss of two protons), and its mass number decreases by four (because of the loss of four nuclear particles). We can represent the emission of an alpha particle with a chemical equation—for example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{231}_{90}\mathrm{Th} \tag{1.5.1.2}$$

How do we know that a product of the reaction is $^{231}_{90}\mathrm{Th}$? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means we must have the same number of protons and neutrons on both sides of the chemical equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose 4 nuclear particles of the original 235, there are 231 remaining. Thus, we use subtraction to identify the isotope of the thorium atom—in this case, $^{231}_{90}\mathrm{Th}$.

Chemists often use the names *parent isotope* and *daughter isotope* to represent the original atom and the product other than the alpha particle. In the previous example, $^{235}_{92}\mathrm{U}$ is the parent isotope, and $^{231}_{90}\mathrm{Th}$ is the daughter isotope. When one element changes into another in this manner, it undergoes *radioactive decay*.

Example 1.5.1.1

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as $^{222}_{86}\mathrm{Rn}$. We represent the alpha particle as $^{4}_{2}\mathrm{He}$ and use subtraction (222 − 4 = 218 and 86 − 2 = 84) to identify the daughter isotope as an isotope of polonium, $^{218}_{84}\mathrm{Po}$:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{218}_{84}\mathrm{Po}$$

? Exercise ? 1.5.1.1

Write the nuclear equation that represents the radioactive decay of polonium-209 by alpha particle emission and identify the daughter isotope.

Answer

$$^{209}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{205}_{82}\text{Pb}$$

The second major type of radioactive emission is called a beta particle, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1− charge. We can also represent a beta particle as $^{0}_{-1}\text{e}$ or β^-. The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e} \tag{1.5.1.3}$$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the electron is assigned an “atomic number” of 1−, equal to its charge.)

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called gamma rays, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*). For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He} + \gamma \tag{1.5.1.4}$$

This may not seem like much energy, but if 1 mol of radon atoms were to decay, the gamma ray energy would be 49 million kJ!

✓ Example 1.5.1.2

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is $^{12}_{5}\text{B}$ while one of the products is an electron, $^{0}_{-1}\text{e}$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus, the complete nuclear equation is as follows:

$$^{12}_{5}\text{B} \rightarrow {}^{12}_{6}\text{C} + {}^{0}_{-1}\text{e} + \gamma$$

The daughter isotope is $^{12}_{6}\text{C}$.

? Exercise 1.5.1.2

Write the nuclear equation that represents the radioactive decay of iodine-131 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Answer

$$^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}\text{e} + \gamma$$

Alpha, beta, and gamma emissions have different abilities to penetrate matter. The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter. Table 1.5.1.1summarizes the properties of the three main types of radioactive emissions.

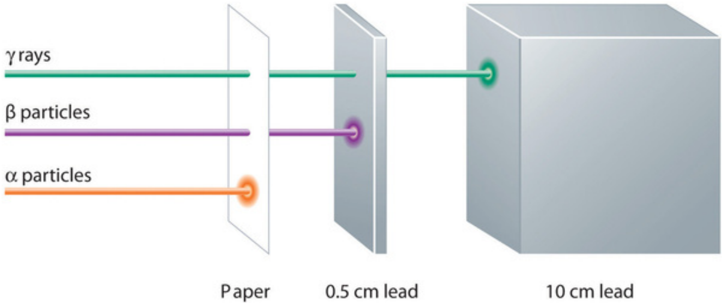


Figure 1.5.1.2: Different emissions exhibit different pentration powers. (CC BY-NC-SA 3.0; anonymous)

Table 1.5.1.1: The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, $^{4}_{2}\text{He}$	β, $^{0}_{-1}\text{e}$	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1−	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called *spontaneous fission* (or fission). Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\text{U} \rightarrow {}^{139}_{56}\text{Ba} + {}^{94}_{36}\text{Kr} + 2\,{}^{1}_{0}\text{n} \tag{1.5.1.5}$$

where $^{1}_{0}\text{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and the mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208.

> *Fission is the radioactive process used in nuclear power plants and one type of nuclear bomb.*

Key Takeaway

The major types of radioactivity include alpha particles, beta particles, and gamma rays.

1.5.2: Nuclear Bombardment Reactions

Positron Emission

There is another type of nuclear emission that results in a very unusual particle, a "positron" or a positively charged electron! This particle is a form of antimatter and almost immediately combines with a normal electron in the environment to produce gamma radiation. During positron emission, a proton in the nucleus converts to a neutron and a positron. The neutron stays in the nucleus but the positron is emitted. Writing the positron as shown below ensures that the atomic numbers and mass numbers are balanced.

$$^{24}_{13}Al \rightarrow ^{24}_{12}Mg + ^{0}_{+1}e \tag{1.5.2.1}$$

Gamma Emission

The previous types of nuclear reaction, alpha emission, beta emission, and positron emission, all involve a single radioactive atom splitting into two pieces. One of the pieces is very small, so you can picture it as a nucleus emitting a small particle. The nucleus changes identity as a result.

Sometimes, a nucleus does not emit a particle at all, but only changes from an energetic unstable state (metastable) to a more stable state while emitting energy in the form of gamma radiation. This type of nuclear equation can be written this way.

$$^{99m}_{43}Tc \rightarrow ^{99}_{43}Tc + ^{0}_{0}\gamma \tag{1.5.2.2}$$

Nuclear Bombardment

Another type of nuclear reaction occurs when scientists purposefully launch two types of nuclei together, or launch a small particle at a nucleus. These reactions have two things written on the left side of the arrow and two or more written on the right.

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{141}_{56}Ba + ^{92}_{36}Kr + 3^{1}_{0}n \tag{1.5.2.3}$$

As in our previous equations, the sum of the atomic numbers on the left has to equal the sum on the right (92 + 0 = 56 + 36 + 0). The sum of the mass numbers on the left has to equal the sum of the mass numbers on the right (235 + 1 = 141 + 92 + 3).

These reactions sometimes result in new heavy elements, and is the way the elements beyond uranium were created. One of the most recent elements to be created this way is Mc:

$$^{48}_{20}Ca + ^{243}_{95}Am \rightarrow ^{291}_{115}Mc \tag{1.5.2.4}$$

Exercise 1.5.2.1

What is the missing isotope in the reaction below?

$$^{241}_{95}Am + ^{4}_{2}He \rightarrow ? + 2^{1}_{0}n \tag{1.5.2.5}$$

Answer

$^{243}_{97}Bk$

1.5.3: Half-Life

Learning Objectives

- To define *half-life*.
- To determine the amount of radioactive substance remaining after a given number of half-lives.

Whether or not a given isotope is radioactive is a characteristic of that particular isotope. Some isotopes are stable indefinitely, while others are radioactive and decay through a characteristic form of emission. As time passes, less and less of the radioactive isotope will be present, and the level of radioactivity decreases. An interesting and useful aspect of radioactive decay is the **half-life**. The half-life of a radioactive isotope is the amount of time it takes for one-half of the radioactive isotope to decay. The half-life of a specific radioactive isotope is constant; it is unaffected by conditions and is independent of the initial amount of that isotope.

Consider the following example. Suppose we have 100.0 g of ^{3}H (tritium, a radioactive isotope of hydrogen). It has a half-life of 12.3 y. After 12.3 y, half of the sample will have decayed to ^{3}He by emitting a beta particle, so that only 50.0 g of the original ^{3}H remains. After another 12.3 y—making a total of 24.6 y—another half of the remaining ^{3}H will have decayed, leaving 25.0 g of ^{3}H. After another 12.3 y—now a total of 36.9 y—another half of the remaining ^{3}H will have decayed, leaving 12.5 g of ^{3}H. This sequence of events is illustrated in Figure $1.5.3.1$.

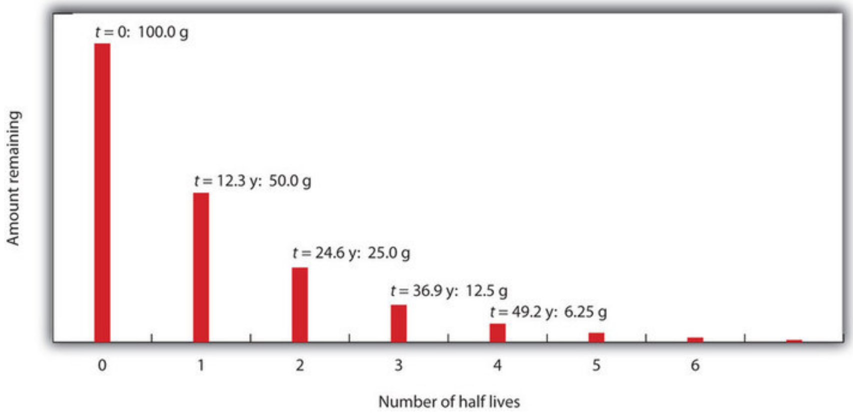


Figure $1.5.3.1$: Radioactive Decay. During each successive half-life, half of the initial amount will radioactively decay.

We can determine the amount of a radioactive isotope remaining after a given number half-lives by using the following expression:

$$\text{amount remaining} = \text{initial amount} \times \left(\frac{1}{2}\right)^{\text{n}} \tag{1.5.3.1}$$

where *n* is the number of half-lives. This expression works even if the number of half-lives is not a whole number.

Example $1.5.3.1$

The half-life of ^{20}F is 11.0 s. If a sample initially contains 5.00 g of ^{20}F, how much ^{20}F remains after 44.0 s?

Solution

If we compare the time that has passed to the isotope's half-life, we note that 44.0 s is exactly 4 half-lives, so we can use Equation ??? with $n = 4$. Substituting and solving results in the following:

$$\begin{aligned} \text{amount remaining} &= 5.00\ g \times \left(\frac{1}{2}\right)^4 \\ &= 5.00\ g \times \frac{1}{16} \\ &= 0.313\ g \end{aligned}$$

Less than one-third of a gram of ^{20}F remains.

Exercise 1.5.3.2

The half-life of ^{44}Ti is 60.0 y. A sample initially contains 0.600 g of ^{44}Ti. How much ^{44}Ti remains after 180.0 y?

Answer

0.075 g.

Half-lives of isotopes range from fractions of a microsecond to billions of years. Table $1.5.3.1$ lists the half-lives of some isotopes.

Table 1.5.3.1: Half-Lives of Various Isotopes

Isotope	Half-Life
^{3}H	12.3 y
^{14}C	5,730 y
^{40}K	1.26 × 10^{9} y
^{51}Cr	27.70 d
^{90}Sr	29.1 y
^{131}I	8.04 d
^{222}Rn	3.823 d
^{235}U	7.04 × 10^{8} y
^{238}U	4.47 × 10^{9} y
^{241}Am	432.7 y
^{248}Bk	23.7 h
^{260}Sg	4 ms

Example 1.5.3.2

The isotope I-125 is used in certain laboratory procedures and has a half-life of 59.4 days. If the initial activity of a sample of I-125 is $32,000$ counts per minute (cpm), how much activity will be present in 178.2 days?

Solution

We begin by determining how many half-lives are represented by 178.2 days:

$$\frac{178.2 \text{ days}}{59.4 \text{ days/half-life}} = 3 \text{ half-lives}$$

Then we simply count activity:

$$\begin{aligned} \text{initial activity } (t_0) &= 32,000 \text{ cpm} \\ \text{after one half-life } &= 16,000 \text{ cpm} \\ \text{after two half-lives } &= 8,000 \text{ cpm} \\ \text{after three half-lives} &= 4,000 \text{ cpm} \end{aligned}$$

Be sure to keep in mind that the initial count is at time zero (t_0) and we subtract from that count at the first half-life. The second half-life has an activity of half the previous count (not the initial count).

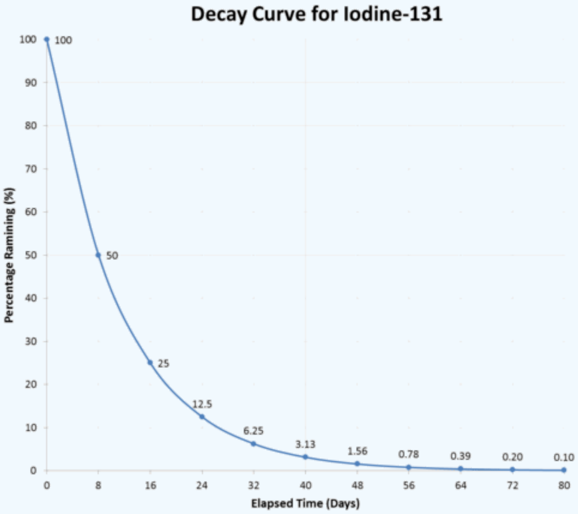
Equation $1.5.3.1$ can be used to calculate the amount of radioactivity remaining after a given time:

$$N_t = N_0 \times (0.5)^{\text{number of half-lives}}$$

where $N_t =$ activity at time t and $N_0 =$ initial activity at time $t = 0$.

If we have an initial activity of $42,000$ cpm, what will the activity be after four half-lives?

$$\begin{aligned} N_t &= N_0(0.5)^4 \\ &= (42,000)\,(0.5)\,(0.5)\,(0.5)\,(0.5) \\ &= 2625 \text{ cpm} \end{aligned}$$



Typical radioactive decay curve.

The graph above illustrates a typical decay curve for $\mathrm{I-125}$. The activity decreases by one-half during each succeeding half-life.

ise

A sample of Ac-225 originally contained 80 grams and after 50 days only 2.5 grams of the original Ac-225 remain. What is the half life of Ac-225?

Answer

10 days

Looking Closer: Half-Lives of Radioactive Elements

Many people think that the half-life of a radioactive element represents the amount of time an element is radioactive. In fact, it is the time required for half—not all—of the element to decay radioactively. Occasionally, however, the daughter element is also radioactive, so its radioactivity must also be considered.

The expected working life of an ionization-type smoke detector (described in the opening essay) is about 10 years. In that time, americium-241, which has a half-life of about 432 y, loses less than 4% of its radioactivity. A half-life of 432 y may seem long to us, but it is not very long as half-lives go. Uranium-238, the most common isotope of uranium, has a half-life of about 4.5 × 10^9 y, while thorium-232 has a half-life of 14 × 10^9 y.

On the other hand, some nuclei have extremely short half-lives, presenting challenges to the scientists who study them. The longest-lived isotope of lawrencium, ^{262}Lr, has a half-life of 3.6 h, while the shortest-lived isotope of lawrencium, ^{252}Lr, has a half-life of 0.36 s. As of this writing, the largest atom ever detected has atomic number 118, mass number 293, and a half-life of 120 ns. Can you imagine how quickly an experiment must be done to determine the properties of elements that exist for so short a time?

Key Takeaways

- Natural radioactive processes are characterized by a half-life, the time it takes for half of the material to decay radioactively.
- The amount of material left over after a certain number of half-lives can be easily calculated.

1.5.3.1: Practice Half-Life

Exercise 1.5.3.1.1

How many years does it take for 4.7 half-lives to pass for a sample of sodium-22? The half-life of sodium-22 is 2.6 years..

Answer

12 years

Exercise 1.5.3.1.1

How many half-lives have passed for a sample of gold-198 that is 3.4 weeks old? The half-life of gold-198 is 2.7 days.

Answer

8.8 half-lives

Exercise 1.5.3.1.1

If you start with 422 mg of calcium-47, how much will remain after 5.00 half-lives have passed?

Answer

13.2 mg

Exercise 1.5.3.1.1

The half-life of sodium-24 is 15.0 hours. If you begin with a sample that has 675 nCi of activity, what activity will remain after 3.75 days? (Hint: how many half-lives have passed for the sample after 3.75 days?)

Hint

6.00 half-lives have passed.

Answer

10.5 nCi of activity remain

1.5.4: Units of Radioactivity

Learning Objectives

- To express amounts of radioactivity in a variety of units.

Previously, we used mass to indicate the amount of radioactive substance present. This is only one of several units used to express amounts of radiation. Some units describe the number of radioactive events occurring per unit time, while others express the amount of a person's exposure to radiation.

Perhaps the direct way of reporting radioactivity is the number of radioactive decays per second. One decay per second is called one becquerel (Bq). Even in a small mass of radioactive material, however, there are many thousands of decays or disintegrations per second. The unit curie (Ci), now defined as 3.7×10^{10} decays per second, was originally defined as the number of decays per second in 1 g of radium. Many radioactive samples have activities that are on the order of microcuries (µCi) or more. Both the becquerel and curie can be used in place of grams to describe quantities of radioactive material. As an example, the amount of americium in an average smoke detector has an activity of 0.9 µCi.

> *The unit becquerel is named after Henri Becquerel, who discovered radioactivity in 1896. The unit curie is named after Polish scientist Marie Curie, who performed some of the initial investigations into radioactive phenomena and discovered the elements, polonium (Po) and radium (Ra) in the early 1900s.*

Example 1.5.4.1

A sample of radium has an activity of 16.0 mCi (millicuries). If the half-life of radium is 1,600 y, how long before the sample's activity is 1.0 mCi?

Solution

The following table shows the activity of the radium sample over multiple half-lives:

Solutions to Example 11.3.1

Time in Years	Activity
0	16.0 mCi
1,600	8.0 mCi
3,200	4.0 mCi
4,800	2.0 mCi
6,400	1.0 mCi

Over a period of 4 half-lives, the activity of the radium will be halved four times, at which point its activity will be 1.0 mCi. Thus, it takes 4 half-lives, or 4 × 1,600 y = 6,400 y, for the activity to decrease to 1.0 mCi.

Exercise 1.5.4.1

A sample of radon has an activity of 60,000 Bq. If the half-life of radon is 15 h, how long before the sample's activity is 3,750 Bq?

Answer

60 hrs.

Other measures of radioactivity are based on the effects it has on living tissue. Radioactivity can transfer energy to tissues in two ways: through the kinetic energy of the particles hitting the tissue and through the electromagnetic energy of the gamma rays being absorbed by the tissue. Either way, the transferred energy—like thermal energy from boiling water—can damage the tissue.

The rad (an acronym for radiation absorbed dose) is a unit equivalent to a gram of tissue absorbing 0.01 J:

$$1 \text{ rad} = 0.01 \text{ J/g}$$

Another unit of radiation absorption is the gray (Gy):

$$1 \text{ Gy} = 100 \text{ rad}$$

The rad is more common. To get an idea of the amount of energy this represents, consider that the absorption of 1 rad by 70,000 g of H_2O (approximately the same mass as a 150 lb person) would increase its temperature by only 0.002°C. This may not seem like a lot, but it is enough energy to break about 1×10^{21} molecular C–C bonds in a person's body. That amount of damage would not be desirable.

Predicting the effects of radiation is complicated by the fact that various tissues are affected differently by different types of emissions. To quantify these effects, the unit **rem** (an acronym for roentgen equivalent, man) is defined as

$$\text{rem} = \text{rad} \times \text{RBE}$$

where RBE is the *relative biological effectiveness factor* is a number greater than or equal to 1 that takes into account the type of radioactive emission and sometimes the type of tissue being exposed. For beta particles, RBE factor equals 1. For alpha particles striking most tissues, the factor is 10, but for eye tissue, the factor is 30. Most radioactive emissions that people are exposed to are on the order of a few dozen millirems (mrem) or less; a medical X ray is about 20 mrem. A sievert (Sv) is a related unit and is defined as 100 rem.

What is a person's annual exposure to radioactivity and radiation? Table 1.5.4.1 lists the sources and annual amounts of radiation exposure. It may surprise you to learn that fully 82% of the radioactivity and radiation exposure we receive is from natural sources —sources we cannot avoid. Fully 10% of the exposure comes from our own bodies—largely from ^{14}C and ^{40}K.

Table 1.5.4.1: Average Annual Radiation Exposure (Approximate)

Source	**Amount (mrem)**
radon gas	200
medical sources	53
radioactive atoms in the body naturally	39
terrestrial sources	28
cosmic sources	28
consumer products	10
nuclear energy	0.05
Total	358

> *Flying from New York City to San Francisco adds 5 mrem to your overall radiation exposure because the plane flies above much of the atmosphere, which protects us from most cosmic radiation.*

The actual effects of radioactivity and radiation exposure on a person's health depend on the type of radioactivity, the length of exposure, and the tissues exposed. Table 1.5.4.2 lists the potential threats to health at various amounts of exposure over short periods of time (hours or days).

Table 1.5.4.2: Effects of Short-Term Exposure to Radioactivity and Radiation

Exposure (rem)	Effect
1 (over a full year)	no detectable effect
∼20	increased risk of some cancers

Exposure (rem)	Effect
∼100	damage to bone marrow and other tissues; possible internal bleeding; decrease in white blood cell count
200–300	visible “burns” on skin, nausea, vomiting, and fatigue
>300	loss of white blood cells; hair loss
∼600	death

One of the simplest ways of detecting radioactivity is by using a piece of photographic film embedded in a badge or a pen. On a regular basis, the film is developed and checked for exposure. A comparison of the exposure level of the film with a set of standard exposures indicates the amount of radiation a person was exposed to.

Figure 1.5.4.1: Detecting Radioactivity. A Geiger counter is a common instrument used to detect radioactivity.

Another means of detecting radioactivity is an electrical device called a Geiger counter (Figure 1.5.4.1). It contains a gas-filled chamber with a thin membrane on one end that allows radiation emitted from radioactive nuclei to enter the chamber and knock electrons off atoms of gas (usually argon). The presence of electrons and positively charged ions causes a small current, which is detected by the Geiger counter and converted to a signal on a meter or, commonly, an audio circuit to produce an audible “click.”

Key Takeaway

- Radioactivity can be expressed in a variety of units, including rems, rads, and curies.

1.5.5: Uses of Radioactive Isotopes

Learning Objectives

- To learn some applications of radioactivity.

Radioactive isotopes have a variety of applications. Generally, however, they are useful either because we can detect their radioactivity or we can use the energy they release.

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A tracer is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some tritium-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium, ^{3}H, is a radioactive isotope of hydrogen.)

Tracers can also be used to follow the steps of a complex chemical reaction. After incorporating radioactive atoms into reactant molecules, scientists can track where the atoms go by following their radioactivity. One excellent example of this is the use of radioactive carbon-14 to determine the steps involved in the photosynthesis in plants. We know these steps because researchers followed the progress of the radioactive carbon-14 throughout the process.

Radioactive isotopes are useful for establishing the ages of various objects. The half-life of radioactive isotopes is unaffected by any environmental factors, so the isotope acts like an internal clock. For example, if a rock is analyzed and is found to contain a certain amount of uranium-235 and a certain amount of its daughter isotope, we can conclude that a certain fraction of the original uranium-235 has radioactively decayed. If half of the uranium has decayed, then the rock has an age of one half-life of uranium-235, or about 4.5×10^9 y. Many analyses like this, using a wide variety of isotopes, have indicated that the age of Earth itself is over 4×10^9 y. In another interesting example of radioactive dating, ^{3}H dating has been used to verify the stated vintages of some old fine wines.

Carbon-14 (half-life is 5,370 y) is particularly useful in determining the age of once-living artifacts (e.g., animal or plant matter). A tiny amount of carbon-14 is produced naturally in the upper reaches of the atmosphere, and living things incorporate some of it into their tissues, building up to a constant, although very low, level. Once a living thing dies, however, it no longer acquires carbon-14, and as time passes, the carbon-14 that was in the tissues decays. If a once-living artifact is discovered and analyzed many years after its death, with the remaining carbon-14 compared to the known constant level, an approximate age of the artifact can be determined. Using such methods, scientists determined that the age of the Shroud of Turin (made of linen, which comes from the flax plant, and purported by some to be the burial cloth of Jesus Christ; Figure 1.5.5.1) is about 600–700 y, not 2,000 y as claimed by some. Scientists were also able to use radiocarbon dating to show that the age of a mummified body found in the ice of the Alps was 5,300 y.

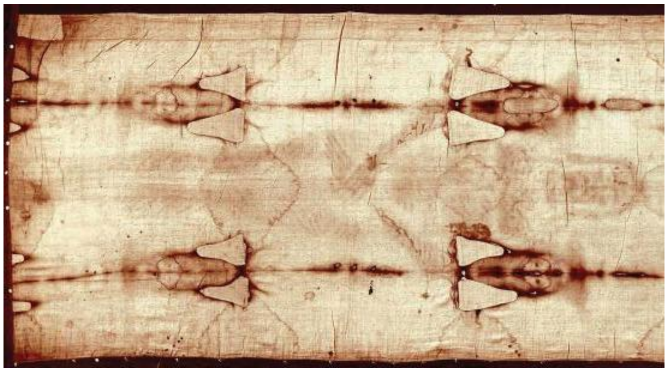


Figure 1.5.5.1: Shroud of Turin. In 1989, several groups of scientists used carbon-14 dating to demonstrate that the age of the Shroud of Turin was only 600–700 y. Many people still cling to a different notion, despite the scientific evidence.

The radiation emitted by some radioactive substances can be used to kill microorganisms on a variety of foodstuffs, which extends the shelf life of these products. Produce such as tomatoes, mushrooms, sprouts, and berries are irradiated with the emissions from coba . This exposure kills a lot of the bacteria that cause spoilage, so the produce stays fresh longer. Eggs and some meat, such as beef, pork, and poultry, can also be irradiated. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.

Radioactive isotopes have numerous medical applications—diagnosing and treating illnesses and diseases. One example of a diagnostic application is using radioactive iodine-131 to test for thyroid activity (Figure 1.5.5.2). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of iodine-131 is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

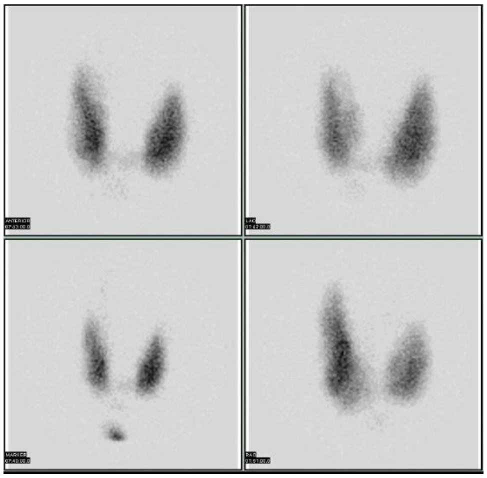


Figure 1.5.5.2: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, Wikipedia, Thyroid Scan(opens in new window) [en.Wikipedia.org].

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect. However, therapeutic applications usually require much larger doses because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor is detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less then 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect but also ease the pain of bone cancers. Table 1.5.5.1lists several radioactive isotopes and their medical uses.

Table 1.5.5.1: Some Radioactive Isotopes That Have Medical Applications

Isotope	Use
^{32}P	cancer detection and treatment, especially in eyes and skin
^{59}Fe	anemia diagnosis
^{60}Co	gamma ray irradiation of tumors
^{99m}Tc	brain, thyroid, liver, bone marrow, lung, heart, and intestinal scanning; blood volume determination
^{131}I	diagnosis and treatment of thyroid function
^{133}Xe	lung imaging
^{198}Au	liver disease diagnosis

In addition to the direct application of radioactive isotopes to diseased tissue, the gamma ray emissions of some isotopes can be directed toward the tissue to be destroyed. Cobalt-60 is a useful isotope for this kind of procedure.

To Your Health: Positron Emission Tomography Scans

One relatively rare form of radioactivity is called *positron emission*. It is similar to beta particle emission, except that instead of emitting an electron, a nucleus emits a positively charged electron, called a *positron*. A positron is actually a piece of antimatter; therefore, when a positron encounters an electron, both particles are converted into high-energy gamma radiation.

Isotopes that emit positrons can be employed in a medical imaging technique called *positron emission tomography* (*PET*). A patient receives a compound containing a positron-emitting isotope, either intravenously or by ingestion. The radioactive compound travels throughout the body, and the patient is then pushed slowly through a ring of sensors that detect the gamma radiation given off by the annihilation of positrons and electrons. A computer connected to the sensors constructs a three-dimensional image of the interior of part or all of the patient's body, allowing doctors to see organs or tumors or regulate the function of various organs (such as the brain or the heart) to diagnose the medical condition of the patient.

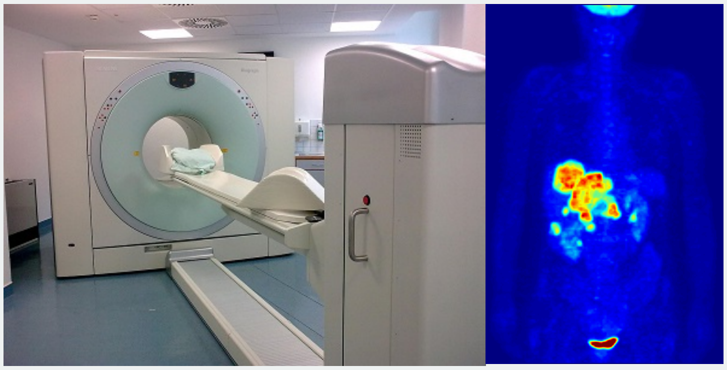


Figure 1.5.5.3: (left) Combined apparatus for positron emission tomography (PET) and X-ray computer tomography (CT), Siemens Biograph (right) Whole-body PET scan using 18F-FDG. Images used with permission from Wikipedia.

Two isotopes that undergo positron emission are carbon-11 and fluorine-18, with half-lives of 20.4 and 110 min, respectively. Both isotopes can be incorporated into sugar molecules and introduced into the body. Doctors can use the intensity of gamma ray emission to find tissues that metabolize the sugar faster than other tissues; fast-metabolizing tissue is one sign of a malignant (i.e., cancerous) tumor. Researchers use similar techniques to map areas of the brain that are most active during specific tasks, such as reading or speaking.

PET is one of many diagnostic and treatment methods that physicians use to improve the quality of our lives. It is one of the many positive uses of radioactivity in society.

Key Takeaway

- Radioactivity has several practical applications, including tracers, medical applications, dating once-living objects, and the preservation of food.

1.5.6: Nuclear Energy

Learning Objectives

- Explain where nuclear energy comes from.
- Describe the difference between fission and fusion.

Nuclear changes occur with a simultaneous release of energy. Where does this energy come from? If we could precisely measure the masses of the reactants and the products of a nuclear reaction, we would notice that the amount of mass drops slightly in the conversion from reactants to products. Consider the following nuclear reaction, in which the molar mass of each species is indicated to four decimal places:

$$\underbrace{{}^{235}\text{U}}_{235.0439} \rightarrow \underbrace{{}^{139}\text{Ba}}_{138.9088} + \underbrace{{}^{94}\text{Kr}}_{93.9343} + \underbrace{2\,{}^{1}\text{n}}_{2\times 1.0087}$$

If we compare the mass of the reactant (235.0439) to the masses of the products (sum = 234.8605), we notice a mass difference of −0.1834 g, or −0.0001834 kg. Where did this mass go?

According to Albert Einstein’s theory of relativity, energy (E) and mass (m) are related by the following equation:

$$E = mc^2 \tag{1.5.6.1}$$

where c is the speed of light, or 3.00×10^8 m/s. In the course of the uranium nuclear chemical reaction, the mass difference is converted to energy, which is given off by the reaction:

$$\begin{aligned} E &= (-0.0001834\ kg)(3.00 \times 10^8\ m/s)^2 \\ &= -1.65 \times 10^{13}\ J \\ &= -1.65 \times 10^{10}\ kJ \end{aligned}$$

That is, 16.5 billion kJ of energy are given off every time 1 mol of uranium-235 undergoes this nuclear reaction. This is an extraordinary amount of energy. Compare it to combustion reactions of hydrocarbons, which give off about 650 kJ/mol of energy for every CH_2 unit in the hydrocarbon—on the order of *hundreds* of kilojoules per mole. Nuclear reactions give off *billions* of kilojoules per mole.

If this energy could be properly harvested, it would be a significant source of energy for our society. Nuclear energy involves the controlled harvesting of energy from fission reactions. The reaction can be controlled because the fission of uranium-235 (and a few other isotopes, such as plutonium-239) can be artificially initiated by injecting a neutron into a uranium nucleus. The overall nuclear equation, with energy included as a product, is then as follows:

$${}^{235}\text{U} + {}^{1}\text{n} \longrightarrow {}^{139}\text{Ba} + {}^{94}\text{Kr} + 3\,{}^{1}\text{n} + \text{energy}$$

Thus, by the careful addition of extra neutrons into a sample of uranium, we can control the fission process and obtain energy that can be used for other purposes.

The Curie Family

Artificial or induced radioactivity was first demonstrated in 1934 by Irène Joliot-Curie and Frédéric Joliot, the daughter and son-in-law of Marie Curie.

Example 1.5.6.1

Plutonium-239 can absorb a neutron and undergo a fission reaction to produce an atom of gold-204 and an atom of phosphorus-31. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Solution

Using the data given, we can write the following initial equation:

$$^1_0\text{n} + ^{239}_{94}\text{Pu} \rightarrow ^{204}_{79}\text{Au} + ^{31}_{15}\text{P} + ?^1_0\text{n}$$

In balanced nuclear equations, the sums of the subscripts on each sides of the equation are the same, as are the sums of the superscripts. The subscripts are already balanced: 0 + 94 = 94 and 79 + 15 = 94. The superscripts on the left equal 240 (1 + 239) but equal 235 (204 + 31) on the right. We need five more mass number units on the right. Five neutrons should be the products of the process for the mass numbers to balance. (Because the atomic number of a neutron is zero, including five neutrons on the right does not change the overall sum of the subscripts.) Thus, the balanced nuclear equation is as follows:

$$^1_0\text{n} + ^{239}_{94}\text{Pu} \rightarrow ^{204}_{79}\text{Au} + ^{31}_{15}\text{P} + 5^1_0\text{n}$$

We predict that the overall process will give off five neutrons.

? Exercise 1.5.6.1

Uranium-238 can absorb a neutron and undergo a fission reaction to produce an atom of cesium-135 and an atom of rubidium-96. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Answer

$$^1_0\text{n} + ^{238}_{92}\text{U} \rightarrow ^{135}_{55}\text{Cs} + ^{96}_{37}\text{Rb} + 8^1_0\text{n}$$

We predict that the overall process will give off eight neutrons.

A nuclear reactor is an apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes. Figure 1.5.6.1 shows a simplified diagram of a nuclear reactor. The energy from the controlled nuclear reaction converts liquid water into high-pressure steam, which is used to run turbines that generate electricity.

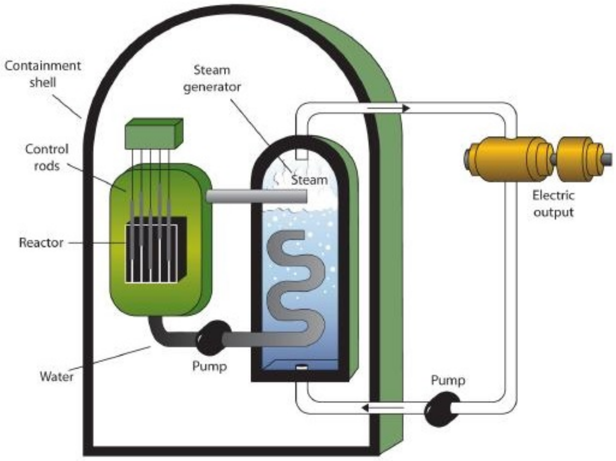


Figure 1.5.6.1: A Diagram of a Nuclear Power Plant for Generating Electricity. The two main components of the power plant are the nuclear reactor itself and the steam-driven turbine and electricity generator.

Notice that the fission of uranium produces two more free neutrons than were present to begin with. These neutrons can themselves stimulate other uranium nuclei to undergo fission, releasing yet more energy and even more neutrons, which can in turn induce even more uranium fission. A single neutron can thus begin a process that grows exponentially in a phenomenon called a chain reaction:

$$1 \rightarrow 2 \rightarrow 4 \rightarrow 8 \rightarrow 16 \rightarrow 32 \rightarrow 64 \rightarrow 128 \rightarrow 256 \rightarrow 512 \rightarrow 1{,}024 \rightarrow 2{,}048 \rightarrow 4{,}096 \rightarrow 8{,}192 \rightarrow 16{,}384 \rightarrow \ldots$$

Because energy is produced with each fission event, energy is also produced exponentially and in an uncontrolled fashion. The quick production of energy creates an explosion. This is the mechanism behind the atomic bomb.

> The first controlled chain reaction was achieved on December 2, 1942, in an experiment supervised by Enrico Fermi in a laboratory underneath the football stadium at the University of Chicago.

Although fairly simple in theory, an atomic bomb is difficult to produce, in part because uranium-235, the isotope that undergoes fission, makes up only 0.7% of natural uranium; the rest is mostly uranium-238, which does not undergo fission. (Remember that the radioactive process that a nucleus undergoes is characteristic of the isotope.) To make uranium useful for nuclear reactors, the uranium in uranium-235 must be *enriched* to about 3%. Enrichment of uranium is a laborious and costly series of physical and chemical separations. To be useful in an atomic bomb, the uranium in uranium-235 must be enriched to 70% or more. At lesser concentrations, the chain reaction cannot sustain itself, so no explosion is produced.

Fusion is another nuclear process that can be used to produce energy. In this process, smaller nuclei are combined to make larger nuclei, with an accompanying release of energy. One example is the hydrogen fusion, which makes helium. While the steps of the process are complicated, the net reaction is:

$$4\,{}^{1}\text{H} \rightarrow {}^{4}\text{He} + 2.58 \times 10^{12}\ J$$

Notice that the amount of energy given off per mole of reactant is only a fraction of the amount given off by the fission of 1 mol of uranium-235. On a mass (per gram) basis, however, the hydrogen fusion emits many times more energy than fission does. In addition, the product of fission is helium gas, not a wide range of isotopes (some of which are also radioactive) produced by fission.

The practical problem is that to perform fusion, extremely high pressures and temperatures are necessary. Currently, the only known stable systems undergoing fusion are the interiors of stars. The conditions necessary for fusion can be created using an atomic bomb, but the resulting fusion is uncontrollable (and the basis for another type of bomb, a hydrogen bomb). Currently, researchers are looking for safe, controlled ways of producing useful energy using fusion.

Career Focus: Nuclear Medicine Technologist

Generally speaking, a radiological technician deals with X ray equipment and procedures. A *nuclear medicine technologist* has similar responsibilities, using compounds containing radioactive isotopes to help diagnose and treat disease.

Nuclear medicine technologists administer the substances containing the radioactive isotope and subsequently operate the apparatus that detects the radiation produced by radioactive decay. The apparatus may be as simple as a piece of photographic film or as complex as a series of computer-controlled electronic detectors. The images obtained by the technologist are interpreted by a specially trained physician.

One of the chief responsibilities of a nuclear medicine technologist is safety. Improper exposure to radioactivity can be harmful to both patient and technologist alike. Therefore, the technologist must adhere to strict safety standards to keep unnecessary exposure as low as possible. The technologist must also know how to dispose of waste materials safely and appropriately.

Key Takeaways

- Nuclear energy comes from tiny mass changes in nuclei as radioactive processes occur.
- In fission, large nuclei break apart and release energy; in fusion, small nuclei merge together and release energy.

1.5.6.1: Practice Nuclear Chemistry

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{241}_{95}Am \rightarrow ? + ^{4}_{2}He \tag{1.5.6.1.1}$$

Answer

$^{237}_{93}Np$ It is alpha emission.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{58}_{28}Ni + ^{1}_{1}H \rightarrow ? + ^{4}_{2}He \tag{1.5.6.1.2}$$

Answer

$^{55}_{27}Co$ It is nuclear bombardment.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{90}_{39}Y \rightarrow ? + ^{0}_{-1}e \tag{1.5.6.1.3}$$

Answer

$^{90}_{40}Zr$ It is beta emission.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{118}_{54}Xe \rightarrow ? + ^{0}_{+1}e \tag{1.5.6.1.4}$$

Answer

$^{118}_{53}I$ It is positron emission.

Exercise 1.5.6.1.1

Two units used to measure absorbed dose of radiation are the rad and the gray (Gy). Their relationship is that 1 Gy = 100 rad, exactly. If a medical treatment involves the absorption of 0.45 krad, what is that in Gy?

Answer

4.5 Gy

Exercise 1.5.6.1.1

The units used to measure radioactivity of a sample are curies (Ci) or bequerel (Bq). Their relationship is that 1 Ci = 3.7 x 10^{10} Bq (2 sig dig relationship). If a radioactive sample is giving off 35 nCi, how many Bq is that?.

Answer

1300 Bq

More Chapter 5 Practice

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{51}_{24}Cr \rightarrow ? + ^{0}_{-1}e \tag{1.5.6.1.5}$$

Answer

$^{51}_{25}Mn$ It is beta emission.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{40}_{18}Ar + ^{4}_{2}He \rightarrow ? + ^{1}_{1}H \tag{1.5.6.1.6}$$

Answer

$^{43}_{19}K$ It is nuclear bombardment.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{214}_{84}Po \rightarrow ? + ^{4}_{2}He \tag{1.5.6.1.7}$$

Answer

$^{210}_{82}Pb$ It is alpha emission.

Exercise 1.5.6.1.1

What is the missing isotope in the reaction below? What type of reaction is it?

$$^{188}_{80}Hg \rightarrow ? + ^{0}_{+1}e \tag{1.5.6.1.8}$$

Answer

$^{188}_{79}Au$ It is positron emission.

Exercise 1.5.6.1.1

Two units used to measure absorbed dose of radiation are the rad and the gray (Gy). Their relationship is that 1 Gy = 100 rad, exactly. If a medical treatment involves the absorption of 75 mrad, what is that in Gy?

Answer

0.00075 Gy

Exercise 1.5.6.1.1

How old (in weeks) is a sample of Ce-141 that has been around for 3.80 half-lives? The half-life of Ce-141 is 32.5 days.

Answer

17.6 weeks

SECTION OVERVIEW

1.6: Chemical Compounds

1.6.1: Ions

Learning Objectives

- Define the two types of ions.
- Use Lewis diagrams to illustrate ion formation.

Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge as a result because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called cations. Most metals become cations when they make ionic compounds.

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called anions. Most nonmetals become anions when they make ionic compounds.

> The names for positive and negative ions are pronounced CAT-eye-ons (cations) and ANN-eye-ons (anions), respectively.

Electron Transfer

We can use electron configurations to illustrate the electron transfer process between sodium atoms and chlorine atoms.

$$\text{Na: } 1s^2 2s^2 2p^6 3s^1$$

As demonstrated here, a sodium atom (Na) has one valence electron in the third principal energy level. It is likely to achieve an octet in its outermost shell by losing its one valence electron. The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The sodium ion, Na^+, has the electron configuration with an octet of electrons from the second principal energy level. It is now the same as that of the noble gas neon. The term **isoelectronic** refers to an atom and an ion of a different atom (or two different ions) that have the same electron configuration. The sodium ion is isoelectronic with the neon atom. The equation below illustrates this process.

$$\begin{array}{lcl} \text{Na} & \rightarrow & \text{Na}^+ + \text{e}^- \\ 1s^2\ 2s^2\ 2p^6\ 3s^1 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

Figure $1.6.1.1$is a graphical depiction of this process.

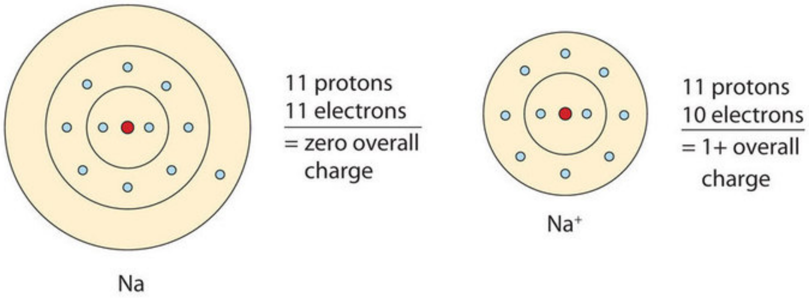


Figure 1.6.1.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Now, let's consider chlorine atom, Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

Only one more electron is needed to achieve an octet in chlorine’s valence shell. When a chlorine atom gains an electron, its outermost principal energy level achieves an octet. In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. This process is illustrated below. (In table salt, this electron comes from the sodium atom.)

$$\begin{array}{lcl} \mathrm{Cl} + \mathrm{e}^- & \rightarrow & \mathrm{Cl}^- \\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5 & & 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 (\text{octet}) \end{array}$$

Figure $1.6.1.2$is a graphical depiction of this process.

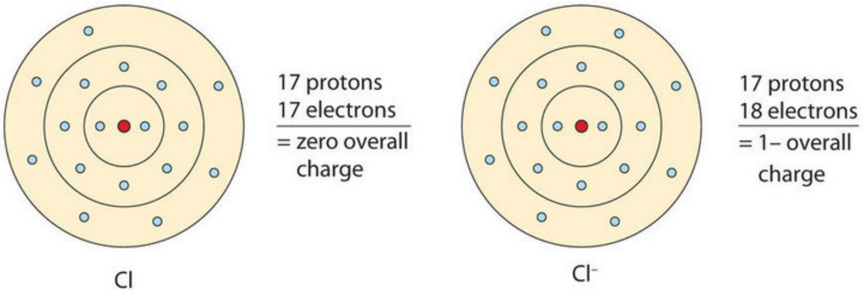


Figure $1.6.1.2$: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge.

With two oppositely charged ions, there is an electrostatic attraction between them because opposite charges attract. The resulting combination is the compound sodium chloride. Notice that there are no leftover electrons. The number of electrons lost by the sodium atom (one) equals the number of electrons gained by the chlorine atom (one), so the compound is electrically neutral. In macroscopic samples of sodium chloride, there are billions and billions of sodium and chloride ions, although there is always the same number of cations and anions.

✓ Example $1.6.1.1$

Write the electron configuration of aluminum atom (Z=13). How many electrons must Al lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Solution

The electron configuration of Al atom is $1s^22s^22p^63s^23p^1$. The second shell has octet ($2s^22p^6$) while the valence shell has 3 electrons ($3s^23p^1$). Mg can achieve octet by losing the 3 valence electrons. The resulting cation is Al^{3+} with electron configuration, $1s^22s^22p^6$.

ise

Write the electron configuration of oxygen atom (Z=8). How many electrons must O lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Answer

The electron configuration of O atom is $1s^22s^22p^4$. The second shell has six electrons ($2s^22p^4$) and needs two electrons to achieve octet. Oxygen will gain 2 electrons. The resulting anion is O^{2-} with electron configuration, $1s^22s^22p^6$.

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge. Figure $1.6.1.3$shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a multiply charged ion. The barium cation is written Ba^{2+}, not Ba^{+2}.

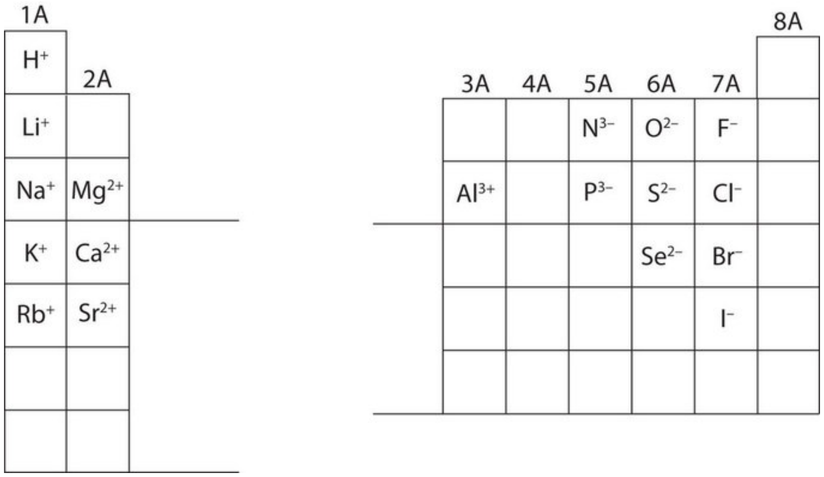


Figure 1.6.1.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

✓ Example 1.6.1.2

Which of these ions is *not likely* to form?

a. Mg^+
b. K^+

Solution

(a) Mg is in Group 2A and has two valence electrons. It achieves octet by losing two electrons to form Mg^{2+} cation. Losing only one electron to form Mg^+ does not make an octet, hence, Mg^+ is not likely to form.

ise

Which of these ions is *not likely* to form?

a. S^{3-}
b. N^{3-}

Answer

(a) S is in Group 6A and has six valence electrons. It achieves octet by gaining two electrons to form S^{2-} anion. Gaining three electrons to form S^{3-}does not make it octet, hence, S^{3-} is not likely to form.

Lewis Diagrams

Chemists use simple diagrams to show an atom's valence electrons and how they transfer. These diagrams have two advantages over the electron shell diagrams. First, they show only valence electrons. Second, instead of having a circle around the chemical symbol to represent the electron shell, they have up to eight dots around the symbol; each dot represents a valence electron. These dots are arranged to the right and left and above and below the symbol, **with no more than two dots on a side**. For example, the representation for sodium is as follows:

and the representation for chlorine is as follows:

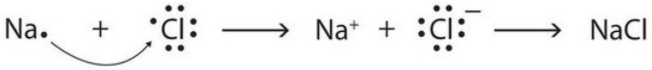
> For the above diagrams, it does not matter what sides the dots are placed on in Lewis diagrams as long as each side has a maximum of two dots.

These diagrams are called **Lewis electron dot diagrams**, or simply Lewis diagrams, after Gilbert N. Lewis, the American chemist who introduced them. To write an element's Lewis dot symbol, place the dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. In other words, **place the dots singly on each side before pairing them**. The Lewis electron dot diagram of fluorine, for example, with seven valence electrons, is constructed as follows:

Figure 1.6.1.4 shows the electron configurations and Lewis diagrams of the elements lithium through neon, which is the entire second period of the periodic table. For the **main group** elements, the **number of valence electrons** is the **same as the group number** listed at the top of the periodic table.

Figure 1.6.1.4: Lewis Electron Dot Diagrams of the Elements Lithium through Neon

The transfer of electrons can be illustrated easily with Lewis diagrams:



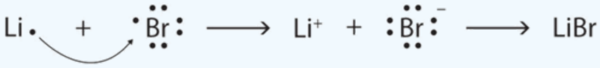
In representing the final formula, the dots are omitted.

Example 1.6.1.3

Starting with lithium and bromine atoms, use Lewis diagrams to show the formation of the ionic compound LiBr.

Solution

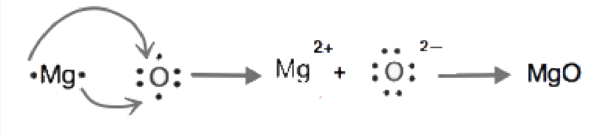
From the periodic table, we see that lithium is in the same column as sodium, so it will have the same valence shell electron configuration. That means that the neutral lithium atom will have the same Lewis diagram that the sodium atom has. Similarly, bromine is in the same column as chlorine, so it will have the same Lewis diagram that chlorine has. Therefore,



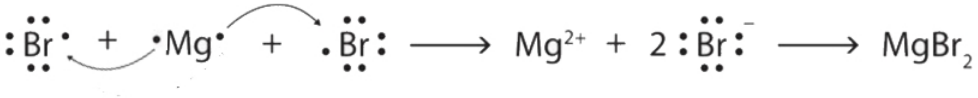
Exercise 1.6.1.3

Starting with magnesium and oxygen atoms, use Lewis diagrams to show the formation of the ionic compound MgO.

Answer



Some ionic compounds have different numbers of cations and anions. In those cases, electron transfer occurs between more than one atom. For example, here is the formation of $MgBr_2$:



Notice that in this example there are two bromide ions (1– charge) needed for every one magnesium ion (2+ charge) in order for the overall charge of the compound to equal zero. This is called **charge balance**. The number of each type of ion is indicated in the formula by the subscript.

Most of the elements that make ionic compounds form an ion that has a characteristic charge. For example, sodium makes ionic compounds in which the sodium ion always has a 1+ charge. Chlorine makes ionic compounds in which the chloride ion always has a 1− charge. Some elements, especially transition metals, can form ions of multiple charges. Figure 1.6.1.5 shows the characteristic charges for some of these ions. As we saw in Figure 1.6.1.1, there is a pattern to the charges on many of the main group ions, but there is no simple pattern for transition metal ions (or for the larger main group elements).

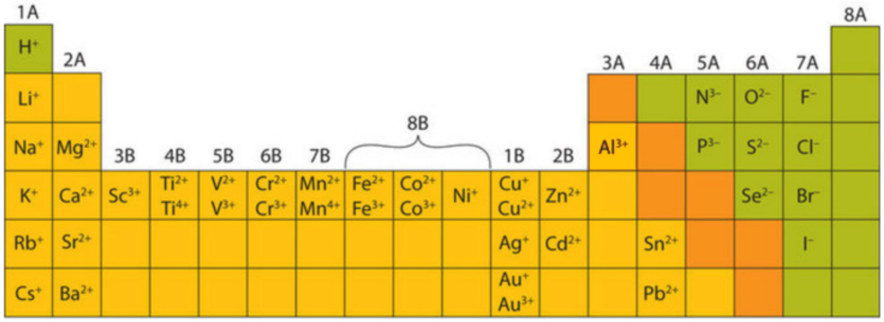


Figure 1.6.1.5: Charges of the Monatomic Ions. Note that some atoms commonly form ions of variable charges.

Key Takeaways

- Ions can be positively charged or negatively charged.
- A Lewis diagram is used to show how electrons are transferred to make ions and ionic compounds.

Exercises

1. What are the two types of ions?
2. Use Lewis diagrams to illustrate the formation of an ionic compound from a potassium atom and an iodine atom.

3. When the following atoms become ions, what charges do they acquire?

a. Li
b. S
c. Ca
d. F

4. Identify each as a cation, an anion, or neither.

a. H^+
b. Cl^-
c. O_2
d. Ba^{2+}
e. CH_4
f. CS_2

5. Identify each as a cation, an anion, or neither.

a. NH_3
b. Br^-
c. H^-
d. Hg^{2+}
e. CCl_4
f. SO_3

6. Write the electron configuration for each ion.

a. Li^+
b. Mg^{2+}

c. F^-
d. S^{2-}

7. Write the electron configuration for each ion.

a. Na^+
b. Be^{2+}
c. Cl^-
d. O^{2-}

8. Draw Lewis diagrams for the ions listed in Exercise 6. Also include Lewis diagrams for the respective neutral atoms as a comparison.

9. Draw Lewis diagrams for the ions listed in Exercise 7. Also include Lewis diagrams for the respective neutral atoms as a comparison.

10. Using Lewis diagrams, show the electron transfer for the formation of LiF.

11. Using Lewis diagrams, show the electron transfer for the formation of MgO.

12. Using Lewis diagrams, show the electron transfer for the formation of Li_2O.

13. Using Lewis diagrams, show the electron transfer for the formation of CaF_2.

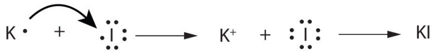
14. What characteristic charge do atoms in the first column of the periodic table have when they become ions?

15. What characteristic charge do atoms in the second column of the periodic table have when they become ions?

16. What characteristic charge do atoms in the third-to-last column of the periodic table have when they become ions?

17. What characteristic charge do atoms in the next-to-last column of the periodic table have when they become ions?

Answers

1. Cations (positive charged) and anions (negative charged).
2. 

3.
    a. 1+
    b. 2−
    c. 2+
    d. 1−

4.
    a. cation
    b. anion
    c. neither
    d. cation
    e. neither
    f. neither

5.
    a. neither
    b. anion
    c. anion
    d. cation
    e. neither
    f. neither

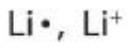
6.
    a. $1s^2$
    b. $1s^22s^22p^6$

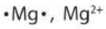
c. $1s^2 2s^2 2p^6$
d. $1s^2 2s^2 2p^6 3s^2 3p^6$

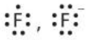
7.

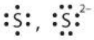
a. $1s^2 2s^2 2p^6$
b. $1s^2$
c. $1s^2 2s^2 2p^6 3s^2 3p^6$
d. $1s^2 2s^2 2p^6$

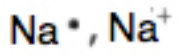
8.

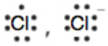
a. 

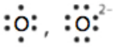
b. 

c. 

d. 

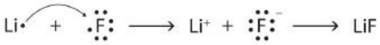




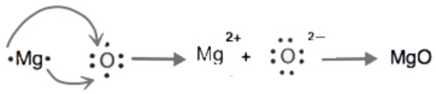


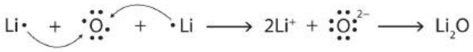
9.

10.



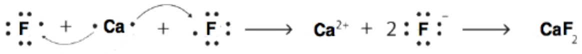
11.





12.

13.



14. 1+

15. 2+

16. 2−

17. 1−

1.6.2: Formulas for Ionic Compounds

Learning Objectives

- Write the chemical formula for a simple ionic compound.
- Recognize polyatomic ions in chemical formulas.

We have already encountered some chemical formulas for simple ionic compounds. A chemical formula is a concise list of the elements in a compound and the ratios of these elements. To better understand what a chemical formula means, we must consider how an ionic compound is constructed from its ions.

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure $1.6.2.1$). As you can see, there are no individual $NaCl$ "particles" in the array; instead, there is a continuous lattice of alternating sodium and chloride ions. However, we can use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers, as a way of describing the compound. In the case of sodium chloride, the ratio of sodium ions to chloride ions, expressed in lowest whole numbers, is 1:1, so we use $NaCl$ (one Na symbol and one Cl symbol) to represent the compound. Thus, $NaCl$ is the chemical formula for sodium chloride, which is a concise way of describing the relative number of different ions in the compound. A macroscopic sample is composed of myriads of NaCl pairs; each individual pair called a **formula unit**. Although it is convenient to think that $NaCl$ crystals are composed of individual $NaCl$ units, Figure $1.6.2.1$ shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.

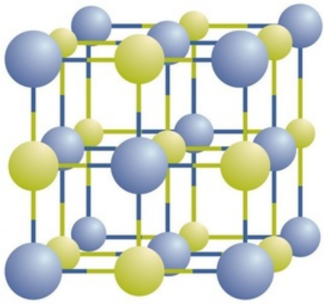


Figure 1.6.2.1: A Sodium Chloride Crystal. A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

The formula for an ionic compound follows several conventions. First, the **cation** is written **before** the **anion**. Because most metals form cations and most nonmetals form anions, formulas typically list the metal first and then the nonmetal. Second, **charges** are **not written** in a formula. Remember that in an ionic compound, the component species are ions, not neutral atoms, even though the formula does not contain charges. Finally, the proper formula for an ionic compound always has ***a net zero charge,*** meaning the total positive charge must equal the total negative charge. To determine the proper formula of any combination of ions, determine how many of each ion is needed to balance the total positive and negative charges in the compound.

> *This rule is ultimately based on the fact that matter is, overall, electrically neutral.*

> By convention, assume that there is only one atom if a subscript is not present. We do not use 1 as a subscript.

If we look at the ionic compound consisting of lithium ions and bromide ions, we see that the lithium ion has a 1+ charge and the bromide ion has a 1− charge. Only one ion of each is needed to balance these charges. The formula for lithium bromide is $LiBr$.

When an ionic compound is formed from magnesium and oxygen, the magnesium ion has a 2+ charge, and the oxygen atom has a 2− charge. Although both of these ions have higher charges than the ions in lithium bromide, they still balance each other in a one-to-one ratio. Therefore, the proper formula for this ionic compound is MgO.

Now consider the ionic compound formed by magnesium and chlorine. A magnesium ion has a 2+ charge, while a chlorine ion has a 1− charge:

$$Mg^{2+}Cl^{-}$$

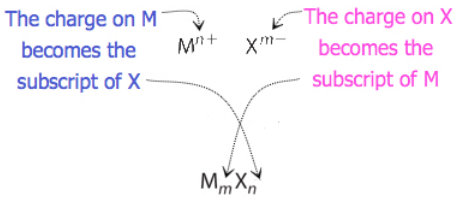
Combining one ion of each does not completely balance the positive and negative charges. The easiest way to balance these charges is to assume the presence of *two* chloride ions for each magnesium ion:

$$Mg^{2+}Cl^{-}Cl^{-}$$

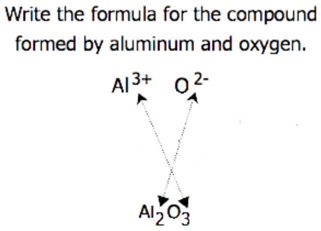
Now the positive and negative charges are balanced. We could write the chemical formula for this ionic compound as $MgClCl$, but the convention is to use a numerical subscript when there is more than one ion of a given type—$MgCl_2$. This chemical formula says that there are one magnesium ion and *two* chloride ions in this formula. (Do not read the "Cl_2" part of the formula as a molecule of the diatomic elemental chlorine. Chlorine does not exist as a diatomic element in this compound. Rather, it exists as two individual chloride ions.) By convention, the **lowest whole number ratio** is used in the formulas of ionic compounds. The formula Mg_2Cl_4 has balanced charges with the ions in a 1:2 ratio, but it is not the lowest whole number ratio.

> By convention, the lowest whole-number ratio of the ions is used in ionic formulas. There are exceptions for certain ions, such as Hg_2^{2+}.

For compounds in which the ratio of ions is not as obvious, the subscripts in the formula can be obtained by **crossing charges**: use the absolute value of the charge on one ion as the subscript for the other ion. This method is shown schematically in Figure 3.3.2.



The image shows M with a charge of n+ and X with a charge of m-. The charge on the M becomes the subscript of X, and the charge on X becomes the subscript of M, making the final product M subscript m X subscript n.



The image asks to write the formula for the compound formed by aluminum and oxygen. It shows Al with a charge of 3+ and O with a charge of 2-. By crossing charges, the final result is Al subscript 2 O subscript 3.

Figure 1.6.2.2: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges.

When crossing charges, it is sometimes necessary to reduce the subscripts to their simplest ratio to write the empirical formula. Consider, for example, the compound formed by Pb^{4+} and O^{2-}. Using the absolute values of the charges on the ions as subscripts gives the formula Pb_2O_4. This simplifies to its correct empirical formula **PbO_2**. The empirical formula has one Pb^{4+} ion and two O^{2-} ions.

✓ Example 1.6.2.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the sodium ion and the sulfur ion
b. the aluminum ion and the fluoride ion
c. the 3+ iron ion and the oxygen ion

Solution

a. To obtain a valence shell octet, sodium forms an ion with a 1+ charge, while the sulfur ion has a 2− charge. Two sodium 1+ ions are needed to balance the 2− charge on the sulfur ion. Rather than writing the formula as NaNaS, we shorten it by convention to $\mathrm{Na_2S}$.
b. The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1− charge. Three fluorine 1− ions are needed to balance the 3+ charge on the aluminum ion. This combination is written as $\mathrm{AlF_3}$.
c. Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxygen atom has a 2− charge as an ion. To balance the positive and negative charges, we look to the least common multiple—6: two iron 3+ ions will give 6+, while three 2− oxygen ions will give 6−, thereby balancing the overall positive and negative charges. Thus, the formula for this ionic compound is $\mathrm{Fe_2O_3}$. Alternatively, use the crossing charges method shown in Figure 3.3.2.

? Exercise 1.6.2.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

Polyatomic Ions

Some ions consist of groups of atoms ***covalently*** bonded together and have an **overall electric charge**. Because these ions contain more than one atom, they are called polyatomic ions. The Lewis structures, names and formulas of some polyatomic ions are found in Table 3.3.1.

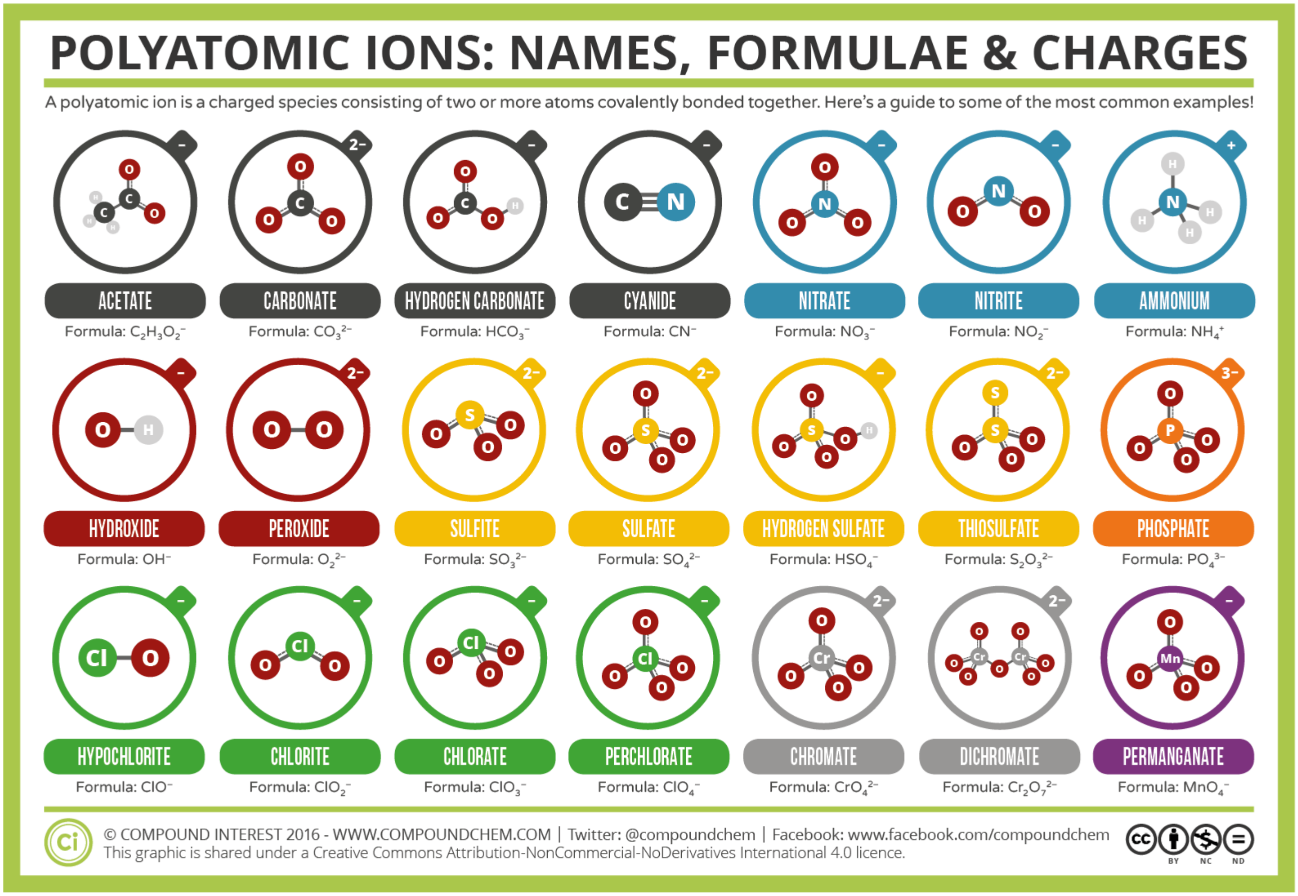


Table 1.6.2.1: Some Polyatomic Ions

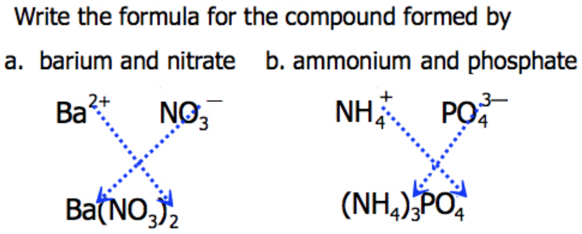
Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table 1.6.2.2lists the ion names and ion formulas of the most common polyatomic ions. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Figure 1.6.2.2lists the most common polyatomic ions.

Table 1.6.2.2: Ion Names and Ion Formulas of Common Polyatomic Ions

Ion Name	Ion Formula
ammonium ion	NH_4^{+1}
hydroxide ion	OH^{-1}
cyanide ion	CN^{-1}
carbonate ion	CO_3^{-2}
bicarbonate or hydrogen carbonate	HCO_3^-
acetate ion	$C_2H_3O_2^{-1}$ **or** $CH_3CO_2^{-1}$
nitrate ion	NO_3^{-1}
nitrite ion	NO_2^{-1}
sulfate ion	SO_4^{-2}
sulfite ion	SO_3^{-2}
phosphate ion	PO_4^{-3}
phosphite ion	PO_3^{-3}

Note that only one polyatomic ion in this Table, the ammonium ion (NH_4^{+1}), is a cation. This polyatomic ion contains one nitrogen and four hydrogens that collectively bear a +1 charge. The remaining polyatomic ions are all negatively-charged and, therefore, are classified as anions. However, only two of these, the hydroxide ion and the cyanide ion, are named using the "-ide" suffix that is typically indicative of negatively-charged particles. The remaining polyatomic anions, which all contain oxygen, in combination with another non-metal, exist as part of a series in which the number of oxygens within the polyatomic unit can vary. As has been repeatedly emphasized in several sections of this text, no two chemical formulas should share a common chemical name. A single suffix, "-ide," is insufficient for distinguishing the names of the anions in a related polyatomic series. Therefore, "-ate" and "-ite" suffixes are employed, in order to denote that the corresponding polyatomic ions are part of a series. Additionally, these suffixes also indicate the relative number of oxygens that are contained within the polyatomic ions. Note that all of the polyatomic ions whose names end in "-ate" contain one more oxygen than those polyatomic anions whose names end in "-ite." Unfortunately, much like the common system for naming transition metals, these suffixes only indicate the *relative* number of oxygens that are contained within the polyatomic ions. For example, the nitrate ion, which is symbolized as NO_3^{-1}, has one more oxygen than the nitrite ion, which is symbolized as NO_2^{-1}. However, the sulfate ion is symbolized as SO_4^{-2}. While both the nitrate ion and the sulfate ion share an "-ate" suffix, the former contains three oxygens, but the latter contains four. Additionally, both the nitrate ion and the sulfite ion contain three oxygens, but these polyatomic ions do not share a common suffix. Unfortunately, the relative nature of these suffixes mandates that the ion formula/ion name combinations of the polyatomic ions must simply be memorized.

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. Two examples are shown below:



The image asks a two-part question: Write the formula for the compound formed by (a) barium and nitrate, and (b) ammonium and phosphate. For (a), Ba has a charge of 2+ and NO_3 has a charge of -1. By crossing charges, the formula is $Ba(NO_3)_2$. For (b), NH_4 has a charge of +1 and PO_4 has a charge of 3-, so by crossing charges the formula is $(NH_4)_3PO_4$.

✓ Example 1.6.2.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the potassium ion and the sulfate ion
b. the calcium ion and the nitrate ion

Solution

a. Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.
b. Calcium ions have a charge of 2+, while nitrate ions have a charge of 1−. We will need two nitrate ions to balance the charge on each calcium ion. The formula for nitrate must be enclosed in parentheses. Thus, we write $Ca(NO_3)_2$ as the formula for this ionic compound.

? Exercise 1.6.2.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

Mg^{2+} and CO_3^{2-} = $MgCO_3$

Answer b:

Al^{3+} and $C_2H_3O_2^-$ = $Al(C_2H_3O_2)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between **metal and nonmetal** elements are usually **ionic**. For example, $CaBr_2$ contains a metallic element (calcium, a group 2A metal) and a nonmetallic element (bromine, a group 7A nonmetal). Therefore, it is most likely an ionic compound. (In fact, it *is* ionic.) In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 5A, and oxygen, from group 6A). It is not an ionic compound; it belongs to the category of covalent compounds discuss elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a **polyatomic ion** in a compound, the compound is **ionic**. For example, if you see the formula $\mathrm{Ba(NO_3)_2}$, you may recognize the "NO_3" part as the nitrate ion, $\mathrm{NO_3^-}$. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the $\mathrm{Ba^{2+}}$ ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

✓ Example 1.6.2.3

Identify each compound as ionic or not ionic.

a. $\mathrm{Na_2O}$
b. $\mathrm{PCl_3}$
c. $\mathrm{NH_4Cl}$
d. $\mathrm{OF_2}$

Solution

a. Sodium is a metal, and oxygen is a nonmetal; therefore, $\mathrm{Na_2O}$ is expected to be ionic.
b. Both phosphorus and chlorine are nonmetals. Therefore, $\mathrm{PCl_3}$ is not ionic.
c. The $\mathrm{NH_4}$ in the formula represents the ammonium ion, $\mathrm{NH_4^+}$, which indicates that this compound is ionic.
d. Both oxygen and fluorine are nonmetals. Therefore, $\mathrm{OF_2}$ is not ionic.

? Exercise 1.6.2.3

Identify each compound as ionic or not ionic.

a. $\mathrm{N_2O}$
b. $\mathrm{FeCl_3}$
c. $\mathrm{(NH_4)_3PO_4}$
d. $\mathrm{SOCl_2}$

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Looking Closer: Blood and Seawater

Science has long recognized that blood and seawater have similar compositions. After all, both liquids have ionic compounds dissolved in them. The similarity may be more than mere coincidence; many scientists think that the first forms of life on Earth arose in the oceans. A closer look, however, shows that blood and seawater are quite different. A 0.9% solution of sodium chloride approximates the salt concentration found in blood. In contrast, seawater is principally a 3% sodium chloride solution, over three times the concentration in blood. Here is a comparison of the amounts of ions in blood and seawater:

Table showing a comparison of the amounts of ions in blood and seawater.

Ion	Percent in Seawater	Percent in Blood
Na^+	2.36	0.322
$Cl^−$	1.94	0.366
Mg^{2+}	0.13	0.002
$SO_4{}^{2−}$	0.09	—
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
$HCO_3{}^−$	0.002	0.165
$HPO_4{}^{2−}$, $H_2PO_4{}^−$	—	0.01

Most ions are more abundant in seawater than they are in blood, with some important exceptions. There are far more hydrogen carbonate ions ($\mathrm{HCO_3^-}$) in blood than in seawater. This difference is significant because the hydrogen carbonate ion and some related ions have a crucial role in controlling the acid-base properties of blood. The amount of hydrogen phosphate ions—$\mathrm{HPO_4^{2-}}$ and $\mathrm{H_2PO_4^-}$—in seawater is very low, but they are present in higher amounts in blood, where they also affect acid-base properties. Another notable difference is that blood does not have significant amounts of the sulfate ion ($\mathrm{SO_4^{2-}}$), but this ion is present in seawater.

Key Takeaways

- Proper chemical formulas for ionic compounds balance the total positive charge with the total negative charge.
- Groups of atoms with an overall charge, called polyatomic ions, also exist.

EXERCISES

1. What information is contained in the formula of an ionic compound?

- Why do the chemical formulas for some ionic compounds contain subscripts, while others do not?

3. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Mg^{2+} and $I^−$
   b. Na^+ and $O^{2−}$

4. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Na^+ and $Br^−$
   b. Mg^{2+} and $Br^−$
   c. Mg^{2+} and $S^{2−}$

5. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. K^+ and $Cl^−$
   b. Mg^{2+} and $Cl^−$
   c. Mg^{2+} and $Se^{2−}$

6. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Na^+ and $N^{3−}$

b. Mg^{2+} and N^{3-}
c. Al^{3+} and S^{2-}

7. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Li^{+} and N^{3-}
b. Mg^{2+} and P^{3-}
c. Li^{+} and P^{3-}

8. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Fe^{3+} and Br^{-}
b. Fe^{2+} and Br^{-}
c. Au^{3+} and S^{2-}
d. Au^{+} and S^{2-}

9. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Cr^{3+} and O^{2-}
b. Cr^{2+} and O^{2-}
c. Pb^{2+} and Cl^{-}
d. Pb^{4+} and Cl^{-}

10. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Cr^{3+} and NO_3^{-}
b. Fe^{2+} and PO_4^{3-}
c. Ca^{2+} and CrO_4^{2-}
d. Al^{3+} and OH^{-}

11. Write the chemical formula for the ionic compound formed by each pair of ions.

a. NH_4^{+} and NO_3^{-}
b. H^{+} and $Cr_2O_7^{2-}$
c. Cu^{+} and CO_3^{2-}
d. Na^{+} and HCO_3^{-}

12. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

a. Ba and S
b. Cs and I

13. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

a. K and S
b. Sc and Br

14. Which compounds would you predict to be ionic?

a. Li_2O
b. $(NH_4)_2O$
c. CO_2
d. $FeSO_3$
e. C_6H_6
f. C_2H_6O

15. Which compounds would you predict to be ionic?

a. $Ba(OH)_2$
b. CH_2O
c. NH_2CONH_2
d. $(NH_4)_2CrO_4$

e. C_8H_{18}
f. NH_3

Answers

1. the ratio of each kind of ion in the compound

2. Sometimes more than one ion is needed to balance the charge on the other ion in an ionic compound.

3.
    a. MgI_2
    b. Na_2O

4.
    a. NaBr
    b. $MgBr_2$
    c. MgS

5.
    a. KCl
    b. $MgCl_2$
    c. MgSe

6.
    a. Na_3N
    b. Mg_3N_2
    c. Al_2S_3

7.
    a. Li_3N
    b. Mg_3P_2
    c. Li_3P

8.
    a. $FeBr_3$
    b. $FeBr_2$
    c. Au_2S_3
    d. Au_2S

9.
    a. Cr_2O_3
    b. CrO
    c. $PbCl_2$
    d. $PbCl_4$

10.
    a. $Cr(NO_3)_3$
    b. $Fe_3(PO_4)_2$
    c. $CaCrO_4$
    d. $Al(OH)_3$

11.
    a. NH_4NO_3
    b. $H_2Cr_2O_7$
    c. Cu_2CO_3
    d. $NaHCO_3$

12.

   a. Ba^{2+}, S^{2-}, BaS
   b. Cs^{+}, I^{-}, CsI

13.

   a. K^{+}, S^{2-}, K_2S
   b. Sc^{3+}, Br^{-}, $ScBr_3$

14.

   a. ionic
   b. ionic
   c. not ionic
   d. ionic
   e. not ionic
   f. not ionic

15.

   a. ionic
   b. not ionic
   c. not ionic
   d. ionic
   e. not ionic
   f. not ionic

1.6.3: Ionic Nomenclature

Learning Objectives

- To use the rules for naming ionic compounds

After learning a few more details about the names of individual ions, you will be a step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges. Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the ***Stock system***, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table $1.6.3.1$: The Common System of Cation Names

Element	Stem	Charge	Name
iron	ferr-	2+	ferrous ion
		3+	ferric ion
copper	cupr-	1+	cuprous ion
		2+	cupric ion
tin	stann-	2+	stannous ion
		4+	stannic ion
lead	plumb-	2+	plumbous ion
		4+	plumbic ion
chromium	chrom-	2+	chromous ion
		3+	chromic ion
gold	aur-	1+	aurous ion
		3+	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table $1.6.3.1$ lists the elements that use the common system, along with their respective cation names.

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, $Cl^−$ is "chlor-" + "-ide ion," or the chloride ion. Similarly, $O^{2−}$ is the oxide ion, $Se^{2−}$ is the selenide ion, and so forth. Table $1.6.3.2$lists the names of some common monatomic ions.

Table $1.6.3.2$: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The polyatomic ions have their own characteristic names, as discussed earlier.

✓ Example 1.6.3.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. $SO_3{}^{2-}$
d. $NH_4{}^+$
e. Cu^+

Answer a

the calcium ion

Answer b

the sulfide ion (from Table 1.6.3.2)

Answer c

the sulfite ion

Answer d

the ammonium ion

Answer e

the copper(I) ion or the cuprous ion (copper can form cations with either a 1+ or 2+ charge, so we have to specify which charge this ion has

? Exercise 1.6.3.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. $SO_4{}^{2-}$
d. Ba^{2+}
e. $HCO_3{}^-$

Answer a

the iron (II) or ferrous ion

Answer b

the iron (III) or ferric ion

Answer c

the sulfate ion

Answer d

the barium ion

Answer e

the bicarbonate ion or hydrogen carbonate ion

✓ Example 1.6.3.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Answer a

Br^-

Answer b

PO_4^{3-}

Answer c

Cu^{2+}

Answer d

Mg^{2+}

? Exercise 1.6.3.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the ferrous ion
d. the potassium ion

Answer a

F^-

Answer b

CO_3^{2-}

Answer c

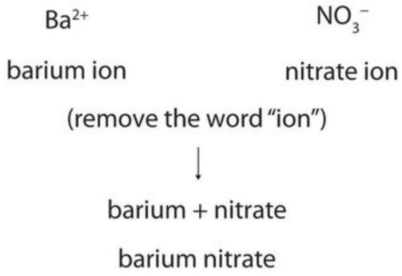
Fe^{2+}

Answer d

K^+

Naming Compounds

Now that we know how to name ions, we are ready to name ionic compounds. We do so by placing the name of the cation first, followed by the name of the anion, and dropping the word *ion* from both parts. For example, what is the name of the compound whose formula is $Ba(NO_3)_2$?



The compound’s name does not indicate that there are two nitrate ions for every barium ion. You must determine the relative numbers of ions by balancing the positive and negative charges.

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride. If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively.

✓ Example 1.6.3.3

Name each ionic compound, using both Stock and common systems if necessary.

a. $Ca_3(PO_4)_2$
b. $(NH_4)_2Cr_2O_7$
c. KCl
d. CuCl
e. SnF_2

Answer a

calcium phosphate

Answer b

ammonium dichromate (the prefix *di-* is part of the name of the anion)

Answer c

potassium chloride

Answer d

copper(I) chloride or cuprous chloride

Answer e

tin(II) fluoride or stannous fluoride

? Exercise 1.6.3.3

Name each ionic compound, using both Stock and common systems if necessary.

a. $ZnBr_2$
b. $Fe(NO_3)_3$
c. Al_2O_3

d. CuF_2
e. AgF

Answer a

zinc bromide

Answer b

iron (III) nitrate or ferric nitrate

Answer c

aluminum oxide

Answer d

copper (II) fluoride or cupric fluoride

Answer e

silver fluoride

Figure 1.6.3.1is a synopsis of how to name simple ionic compounds.

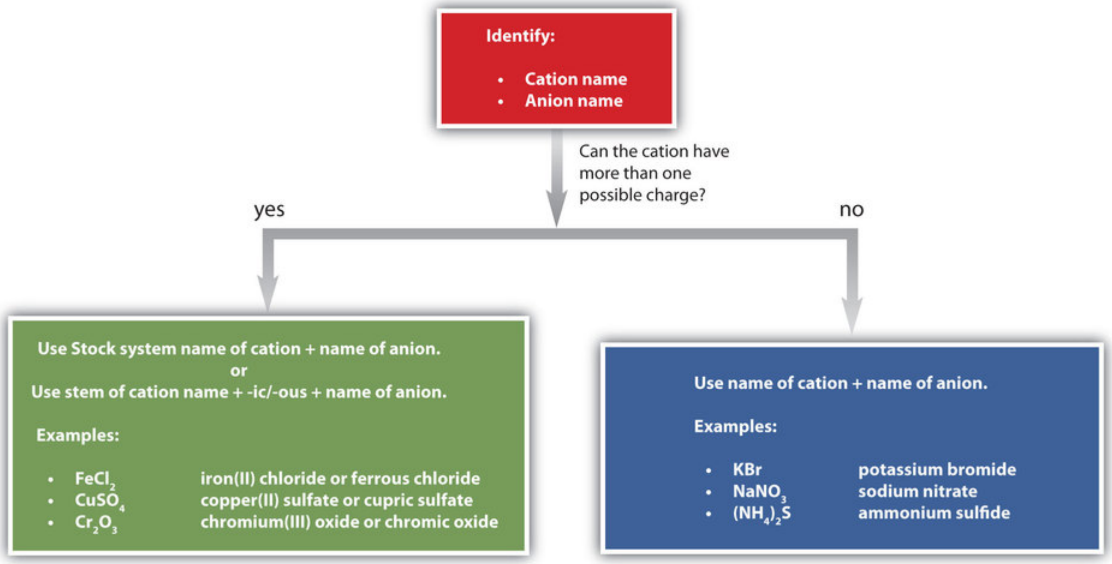


Figure 1.6.3.1: A Guide to Naming Simple Ionic Compounds. Follow these steps to name a simple ionic compound. Identify the cation name and the anion name. If the cation can have more than one possible charge, either use the Stock system name of the cation and name of the anion, or use the stem of the cation name and -ic/-ous and the name of the anion. Examples of this would be $FeCl_2$, which is iron(II) chloride of ferrous chloride, $CuSO_4$, which is copper(II) sulfate or cupric sulfate, and Cr_2O_3, which is chromium(III) oxide or chromic oxide. If the cation can not have more than one possible charge, use the name of the cation and the name of the anion. Examples of this would be KBr, which is potassium bromide, $NaNO_3$, which is sodium nitrate, and $(NH_4)_2S$, which is ammonium sulfide.

KEY TAKEAWAY

- Each ionic compound has its own unique name that comes from the names of the ions.

EXERCISES

1. Briefly describe the process for naming an ionic compound.

- In what order do the names of ions appear in the names of ionic compounds?

3. Which ionic compounds can be named using two different systems? Give an example.

4. Name each ion.

- a. Ra^{2+}
- b. P^{3-}
- c. $H_2PO_4^-$
- d. Sn^{4+}

5. Name each ion.

- a. Cs^+
- b. As^{3-}
- c. HSO_4^-
- d. Sn^{2+}

6. Name the ionic compound formed by each pair of ions.

- a. Na^+ and Br^-
- b. Mg^{2+} and Br^-
- c. Mg^{2+} and S^{2-}

7. Name the ionic compound formed by each pair of ions.

- a. K^+ and Cl^-
- b. Mg^{2+} and Cl^-
- c. Mg^{2+} and Se^{2-}

8. Name the ionic compound formed by each pair of ions.

- a. Na^+ and N^{3-}
- b. Mg^{2+} and N^{3-}
- c. Al^{3+} and S^{2-}

9. Name the ionic compound formed by each pair of ions.

- a. Li^+ and N^{3-}
- b. Mg^{2+} and P^{3-}
- c. Li^+ and P^{3-}

10. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

- a. Fe^{3+} and Br^-
- b. Fe^{2+} and Br^-
- c. Au^{3+} and S^{2-}
- d. Au^+ and S^{2-}

11. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

- a. Cr^{3+} and O^{2-}
- b. Cr^{2+} and O^{2-}
- c. Pb^{2+} and Cl^-
- d. Pb^{4+} and Cl^-

12. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

- a. Cr^{3+} and NO_3^-
- b. Fe^{2+} and PO_4^{3-}
- c. Ca^{2+} and CrO_4^{2-}
- d. Al^{3+} and OH^-

13. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

- a. NH_4^+ and NO_3^-
- b. K^+ and $Cr_2O_7^{2-}$
- c. Cu^+ and CO_3^{2-}
- d. Na^+ and HCO_3^-

14. Give two names for each compound.

    a. $Al(HSO_4)_3$
    b. $Mg(HSO_4)_2$

15. Give two names for each compound.

    a. $Co(HCO_3)_2$
    b. $LiHCO_3$

AnswerS

1. Name the cation and then the anion but don't use numerical prefixes.
2. the cation name followed by the anion name
3. Ionic compounds in which the cation can have more than one possible charge have two naming systems. $FeCl_3$ is either iron(III) chloride or ferric chloride (answers will vary).

4.

    a. the radium ion
    b. the phosphide ion
    c. the dihydrogen phosphate ion
    d. the tin(IV) ion or the stannic ion

5.

    a. the cesium ion
    b. the arsenide ion
    c. the hydrogen sulfate ion
    d. the tin(II) ion or the stannous ion

6.

    a. sodium bromide
    b. magnesium bromide
    c. magnesium sulfide

7.

    a. potassium chloride
    b. magnesium chloride
    c. magnesium selenide

8.

    a. sodium nitride
    b. magnesium nitride
    c. aluminum sulfide

9.

    a. lithium nitride
    b. magnesium phosphide
    c. lithium phosphide

10.

    a. iron(III) bromide or ferric bromide
    b. iron(II) bromide or ferrous bromide
    c. gold(III) sulfide or auric sulfide
    d. gold(I) sulfide or aurous sulfide

11.

    a. chromium(III) oxide or chromic oxide

b. chromium(II) oxide or chromous oxide
c. lead(II) chloride or plumbous chloride
d. lead(IV) chloride or plumbic chloride

12.

a. chromium(III) nitrate or chromic nitrate
b. iron(II) phosphate or ferrous phosphate
c. calcium chromate
d. aluminum hydroxide

13.

a. ammonium nitrate
b. potassium dichromate
c. copper(I) carbonate or cuprous carbonate
d. sodium hydrogen carbonate or sodium bicarbonate

14.

a. aluminum hydrogen sulfate or aluminum bisulfate
b. magnesium hydrogen sulfate or magnesium bisulfate

15.

a. cobalt hydrogen carbonate or cobalt bicarbonate
b. lithium hydrogen carbonate or lithium bicarbonate

1.6.3.1: Practice Naming

Exercise 1.6.3.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

$MgSO_4$

Hint

It is ionic. It is predictable, so no Roman numerals are needed. Just name cation and name anion.

Answer

magnesium sulfate

Exercise 1.6.3.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

Cu_2S

Hint

It is ionic. It is variable, so you must determine the charge needed to make a total of zero charge. Roman numeral is needed.

Answer

copper (I) sulfide   [The copper must be Cu^{1+} in order for two of them to balance the one S^{2-}]

Exercise 1.6.3.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

$Fe(NO_3)_3$

Hint

It is ionic. It is variable, so you must determine the charge needed to make a total of zero charge. Roman numeral is needed.

Answer

iron (III) nitrate

Exercise 1.6.3.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

$Al(ClO_3)_3$

Hint

It is ionic. It is predictable, so no Roman numerals are needed. Just name cation and name anion.

Answer

aluminum chlorate

Exercise 1.6.3.1.1

Name the compound below. (We are not dealing with peroxides. The anion is regular oxide.) Hint: Is it ionic or covalent? Is the cation predictable or variable?

PtO_2

Hint

It is ionic. It is variable, so you must determine the charge needed to make a total of zero charge. Roman numeral is needed.

Answer

platinum (IV) oxide

Exercise 1.6.3.1.1

Write the formula of the compound below. Hint: draw each ion first. Then put them together so they make a compound with zero charge.

calcium nitride

Hint

Ca^{2+} and N^{3-}

Answer

Ca_3N_2

Exercise 1.6.3.1.1

Write the formula of the compound below. Hint: draw each ion first. Then put them together so they make a compound with zero charge.

lithium bisulfite

Hint

Li^{1+} and $HSO_3{}^{1-}$

Answer

$LiHSO_3$

Exercise 1.6.3.1.1

Write the formula of the compound below. Hint: draw each ion first. Then put them together so they make a compound with zero charge.

nickel (III) acetate

Hint

Ni^{3+} and $C_2H_3O_2{}^{1-}$

Answer

$Ni(C_2H_3O_2)_3$

1.6.4: Covalent Compounds - Formulas and Names

Learning Objectives

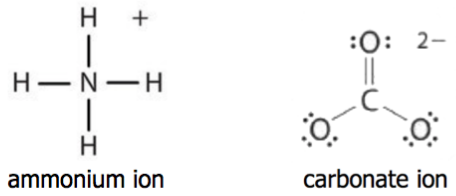
- Identify covalent and ionic compounds.
- Determine the chemical formula of a simple covalent compound from its name.
- Determine the name of a simple covalent compound from its chemical formula.

COVALENT AND IONIC COMPOUNDS

What elements make covalent bonds? Covalent bonds form when two or more nonmetals combine. For example, both hydrogen and oxygen are nonmetals, and when they combine to make water, they do so by forming covalent bonds. Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds.

As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals), silicon dioxide (SiO_2) will be a covalently bound molecule (a semi-metal and a non-metal) and $MgCl_2$ will be ionic (a metal and a non-metal).

A polyatomic ion is an ion composed of two or more atoms that have a charge as a group (poly = many). The ammonium ion (see figure below) consists of one nitrogen atom and four hydrogen atoms. Together, they comprise a single ion with a 1+ charge and a formula of NH_4^+. The carbonate ion (see figure below) consists of one carbon atom and three oxygen atoms and carries an overall charge of 2−. The formula of the carbonate ion is CO_3^{2-}.



The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. Several examples are found in Table 3.3.1. Nonmetal atoms in polyatomic ions are joined by covalent bonds, but the ion as a whole participates in ionic bonding. For example, ammonium chloride (NH_4Cl) has ionic bonding between a polyatomic ion, NH_4^+, and Cl^- ions, but within the ammonium ion (NH_4^+), the nitrogen and hydrogen atoms are connected by covalent bonds (shown above).

Both ionic and covalent bonding are also found in calcium carbonate. Calcium carbonate ($CaCO_3$) has ionic bonding between calcium ion Ca^{2+} and a polyatomic ion, CO_3^{2-}, but within the carbonate ion (CO_3^{2-}), the carbon and oxygen atoms are connected by covalent bonds (shown above).

Characteristics of Covalent (Molecular) Compounds

Compounds that contain covalent bonds (also called molecular compounds) exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds (discussed in Section 3.6). For example, water (molecular compound) boils at 100 °C while sodium chloride (ionic compound) boils at 1413 °C. In fact, many covalent compounds are liquids or gases at room temperature, and, in their solid states, they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good conductors of electricity when dissolved in water, most covalent compounds, being electrically neutral, are poor conductors of electricity in any state. The attraction between molecules (called intermolecular forces) will be discussed in more detail in Section 8.1

Example 1.6.4.1

Is each compound formed from ionic bonds, covalent bonds, or both?

a. Na_2O
b. Na_3PO_4

c. N_2O_4

Answer a

The elements in Na_2O are a metal and a nonmetal, which form ionic bonds.

Answer b

Because sodium is a metal and we recognize the formula for the phosphate ion, we know that this compound is ionic. However, within the polyatomic phosphate ion, the atoms are held together by covalent bonds, so this compound contains both ionic and covalent bonds.

Answer c

The elements in N_2O_4| are both nonmetals, rather than a metal and a nonmetal. Therefore, the atoms form covalent bonds.

Exercise 1.6.4.1

Is each compound are formed from ionic bonds, covalent bonds, or both?

a. $Ba(OH)_2$
b. F_2
c. PCl_3

Answer a:

both

Answer b:

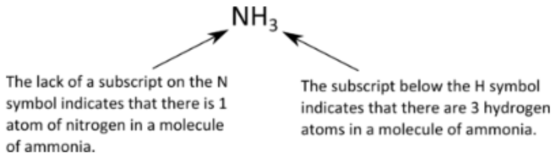
covalent

Answer c:

covalent

MOLECULAR FORMULAS

The chemical formulas for covalent compounds are referred to as **molecular formulas** because these compounds exist as separate, discrete molecules. Typically, a molecular formula begins with the nonmetal that is closest to the lower left corner of the periodic table, except that hydrogen is almost never written first (H_2O is the prominent exception). Then the other nonmetal symbols are listed. Numerical subscripts are used if there is more than one of a particular atom. For example, we have already seen CH_4, the molecular formula for methane. Below is the molecular formula of ammonia, NH_3.



NH_3. An arrow points to the N and says, "The lack of a subscript on the N symbol indicates that there is 1 atom of nitrogen in a molecule of ammonia". An arrow points to H and says, "The subscript below the H symbol indicates that there are 3 hydrogen atoms in a molecule of ammonia".

NAMING COVALENT COMPOUNDS

Naming *binary* (two-element) covalent compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 1.6.4.1 lists these numerical prefixes. Normally, no prefix is added to the first element's name if there is only one atom of the first element in a molecule. If the second element is oxygen, the trailing vowel is usually omitted from the end of a polysyllabic prefix but not a monosyllabic one (that is, we would say "monoxide" rather than "monooxide" and "trioxide" rather than "troxide").

Table 1.6.4.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

*This prefix is not used for the first element's name.

Let us practice by naming the compound whose molecular formula is CCl_4. The name begins with the name of the first element—carbon. The second element, chlor*ine*, becomes chlor*ide*, and we attach the correct numerical prefix ("tetra-") to indicate that the molecule contains four chlorine atoms. Putting these pieces together gives the name *carbon tetrachloride* for this compound.

✓ Example 1.6.4.2

Write the molecular formula for each compound.

a. chlorine trifluoride
b. phosphorus pentachloride
c. sulfur dioxide
d. dinitrogen pentoxide

Solution

If there is no numerical prefix on the first element's name, we can assume that there is only one atom of that element in a molecule.

a. ClF_3
b. PCl_5
c. SO_2
d. N_2O_5 (The *di*- prefix on nitrogen indicates that two nitrogen atoms are present.)

? Exercise 1.6.4.2

Write the molecular formula for each compound.

a. nitrogen dioxide
b. dioxygen difluoride
c. sulfur hexafluoride
d. selenium monoxide

Answer a:

a. NO_2

Answer b:

O_2F_2

Answer c:

SF_6

Answer d:

SeO

Because it is so unreactive, sulfur hexafluoride is used as a spark suppressant in electrical devices such as transformers.

✓ Example 1.6.4.3

Write the name for each compound.

a. BrF_5
b. S_2F_2
c. CO

Solution

a. bromine pentafluoride
b. disulfur difluoride
c. carbon monoxide

? Exercise 1.6.4.3

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system that we will introduce in in a separate section.

Key Takeaways

- The chemical formula of a simple covalent compound can be determined from its name.
- The name of a simple covalent compound can be determined from its chemical formula.

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1.6.5: Acids

Learning Objectives

- Define *acid*.
- Name a simple acid.

There is one other group of compounds that is important to us—acids—and these compounds have interesting chemical properties. Initially, we will define an acid as an ionic compound of the H^+ cation dissolved in water. To indicate that something is dissolved in water, we will use the phase label (aq) next to a chemical formula (where aq stands for "aqueous," a word that describes something dissolved in water). If the formula does not have this label, then the compound is treated as a molecular compound rather than an acid.

Acids have their own nomenclature system. If an acid is composed of only hydrogen and one other element, the name is *hydro-* + the stem of the other element + *-ic acid*. For example, the compound HCl(aq) is hydrochloric acid, while H_2S(aq) is hydrosulfuric acid. If these acids were not dissolved in water, the compounds would be called hydrogen chloride and hydrogen sulfide, respectively. Both of these substances are well known as molecular compounds; when dissolved in water, however, they are treated as acids.

If a compound is composed of hydrogen ions and a polyatomic anion, then the name of the acid is derived from the stem of the polyatomic ion's name. Typically, if the anion name ends in -ate, the name of the acid is the stem of the anion name plus *-ic acid*; if the related anion's name ends in -ite, the name of the corresponding acid is the stem of the anion name plus *-ous acid*. Table 1.6.5.1 lists the formulas and names of a variety of acids that you should be familiar with. You should recognize most of the anions in the formulas of the acids.

Table 1.6.5.1 Names and Formulas of Acids

Formula	Name
$HC_2H_3O_2$	acetic acid
$HClO_3$	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
HF	hydrofluoric acid
HNO_3	nitric acid
$H_2C_2O_4$	oxalic acid
$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid

Note: The "aq" label is omitted for clarity.

Example 1.6.5.1

Name each acid without consulting Table 3.9.

1. HBr
2. H_2SO_4

Solution

1. As a binary acid, the acid's name is *hydro-* + stem name + *-ic acid*. Because this acid contains a bromine atom, the name is hydrobromic acid.
2. Because this acid is derived from the sulfate ion, the name of the acid is the stem of the anion name + *-ic acid*. The name of this acid is sulfuric acid.

Exercise 1.6.5.1

Name each acid.

a. HF
b. HNO_2

Answer a

hydrofluoric acid

Answer b

nitrous acid

All acids have some similar properties. For example, acids have a sour taste; in fact, the sour taste of some of our foods, such as citrus fruits and vinegar, is caused by the presence of acids in food. Many acids react with some metallic elements to form metal ions and elemental hydrogen. Acids make certain plant pigments change colors; indeed, the ripening of some fruits and vegetables is caused by the formation or destruction of excess acid in the plant. In Chapter 12, we will explore the chemical behavior of acids.

Acids are very prevalent in the world around us. We have already mentioned that citrus fruits contain acid; among other compounds, they contain citric acid, $H_3C_6H_5O_7(aq)$. Oxalic acid, $H_2C_2O_4(aq)$, is found in spinach and other green leafy vegetables. Hydrochloric acid not only is found in the stomach (stomach acid), but also can be bought in hardware stores as a cleaner for concrete and masonry. Phosphoric acid is an ingredient in some soft drinks.

Key Takeaways

- An acid is a compound of the H^+ ion dissolved in water.
- Acids have their own naming system.
- Acids have certain chemical properties that distinguish them from other compounds.

1.6.5.1: More Practice Naming

Exercise 1.6.5.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

CuCN

Hint

It is ionic. It is variable, so you must determine the charge needed to make a total of zero charge. Roman numeral is needed.

Answer

copper (I) cyanide

Exercise 1.6.5.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

N_2O_3

Hint

It is covalent. There is no cation. Use prefixes to tell how many of each element are in the formula.

Answer

dinitrogen trioxide

Exercise 1.6.5.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

$Sn_3(PO_4)_2$

Hint

It is ionic. It is variable, so you must determine the charge needed to make a total of zero charge. Roman numeral is needed.

Answer

tin (II) phosphate

Exercise 1.6.5.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

SO_3

Hint

It is covalent. There is no cation. Use prefixes to tell how many of each element are in the formula.

Answer

sulfur trioxide

Exercise 1.6.5.1.1

Name the compound below. Hint: Is it ionic or covalent? Is the cation predictable or variable?

$Ba(OH)_2$

Hint

It is ionic. It is predictable, so no Roman numeral is needed. Just name cation and name anion.

Answer

barium hydroxide

Exercise 1.6.5.1.1

Write the formula of the compound below. Hint: is it ionic or covalent?

cadmium bromide

Hint

It is ionic. Draw ions then put them together to make zero charge.

Answer

$CdBr_2$

Exercise 1.6.5.1.1

Write the formula of the compound below. Hint: is it ionic or covalent?

nitrogen dioxide

Hint

It is covalent. Use prefixes to figure out formula.

Answer

NO_2

Exercise 1.6.5.1.1

Write the formula of the compound below. Hint: is it ionic or covalent?

sulfur hexafluoride

Hint

It is covalent. Use prefixes to figure out formula.

Answer

SF_6

1.6.6: Lewis Structures of Ionic Compounds- Electrons Transferred

Learning Objectives

- State the octet rule.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something: many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the *n* = 2 shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \curvearrowright \cdot\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

resulting in two ions—the Na^+ ion and the Cl^- ion:

$$\mathbf{Na}^+ \qquad \mathbf{:}\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^-$$

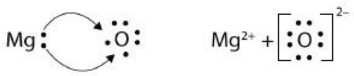
$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Na}^+ + \mathbf{:}\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^- \rightarrow Na^+Cl^- \;\; or \;\; NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

$$\mathbf{Mg}^{2+} + \left[:\underset{\cdot\cdot}{\ddot{\mathbf{O}}}:\right]^{2-} \qquad Mg^{2+}O^{2-} \text{ or } MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na}\cdot \curvearrowright \cdot\underset{\cdot}{\ddot{\mathbf{O}}}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:

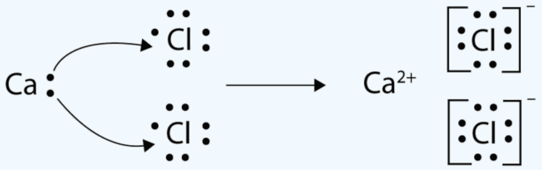
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

✓ Example 1.6.6.1: Synthesis of Calcium Chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:

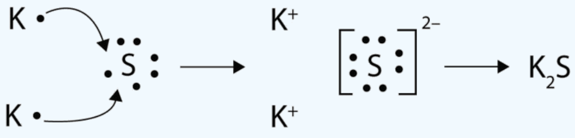


The oppositely charged ions attract each other to make $CaCl_2$.

? Exercise 1.6.6.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

Answer



Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

1.6.6.1: Covalent Lewis Structures- Electrons Shared

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

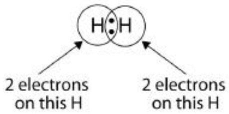
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H}$$

The two H atoms can share their electrons:

$$\mathbf{H:H}$$

We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



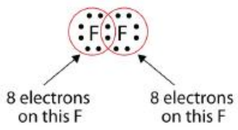
Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:

H–H

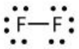
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:

These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.

Note that each F atom has a complete octet around it now:

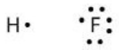


We can also write this using a dash to represent the shared electron pair:

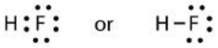


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

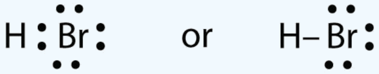
✓ Example 1.6.6.1.1:

Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:

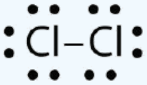
The two atoms can share their unpaired electron:



? Exercise 1.6.6.1.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer



When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far, and there are still some electrons that remain unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of C_2H_4. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two C atoms} = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

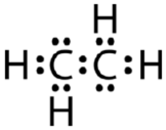


Figure 1.6.6.1.1 : Incorrect dot structure of ethene. (CK12 License)

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

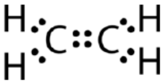


Figure 1.6.6.1.2 : Correct dot structure for ethene. (CK12 License)

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

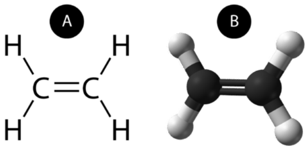


Figure 1.6.6.1.3 : (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

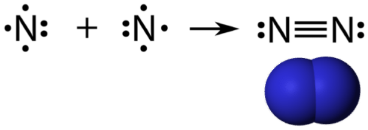


Figure 1.6.6.1.4 : Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributions & Attributions

-
- Anonymous by request

1.6.6.2: Writing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

For CO_3^{2-}, for example, we add two electrons to the total because of the −2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 1.6.6.2.1: Water

Write the Lewis Structure for H_2O.

Solution

Solutions to Example 10.4.1

Steps for Writing Lewis Structures	Example 1.6.6.2.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.

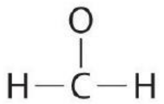
Steps for Writing Lewis Structures	Example 1.6.6.2.1
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

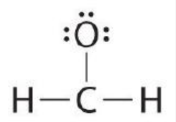
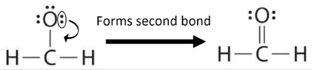
✓ Example 1.6.6.2.2

Write the Lewis structure for the CH_2O molecule

Solution

Solutions to Example 10.4.2

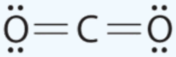
Steps for Writing Lewis Structures	Example 1.6.6.2.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.

Steps for Writing Lewis Structures	Example 1.6.6.2.2
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 1.6.6.2.1

Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO_2



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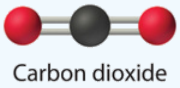
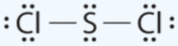


Answer SCl_2



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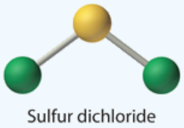


The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

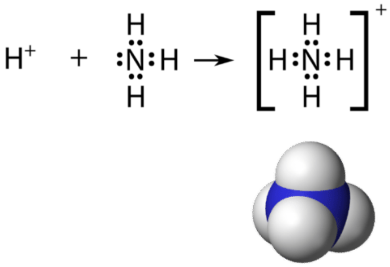


Figure 1.6.6.2.3 : The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$$1 \text{ N atom} = 5 \text{ valence electrons}$$

$$4 \text{ H atoms} = 4 \times 1 = 4 \;\; \text{valence electrons}$$

$$\text{subtract 1 electron for the } 1+ \text{ charge of the ion}$$

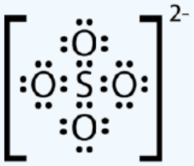
$$\text{total of 8 valence electrons in the ion}$$

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

✓ Exercise 1.6.6.2.2

Draw the Lewis electron dot structure for the sulfate ion.

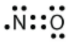
Answer (CK12 License)



Exceptions to the Octet Rule (BC Campus)

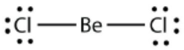
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

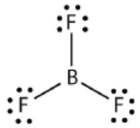


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

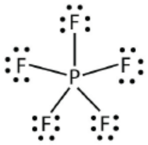
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

✓ Example 1.6.6.2.3: Octet Violations

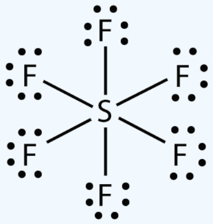
Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:

b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



? Exercise 1.6.6.2.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer

The Xe atom has an expanded valence shell with more than eight electrons around it.

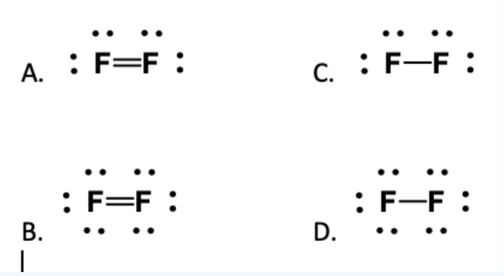
Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

1.6.6.3: Practice Lewis Structures

Exercise 1.6.6.3.1

Which is the Lewis (electron dot) structure of the F_2 molecule?



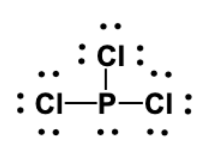
Answer

D

Exercise 1.6.6.3.1

Draw the Lewis (electron dot) structure of PCl_3.

Answer

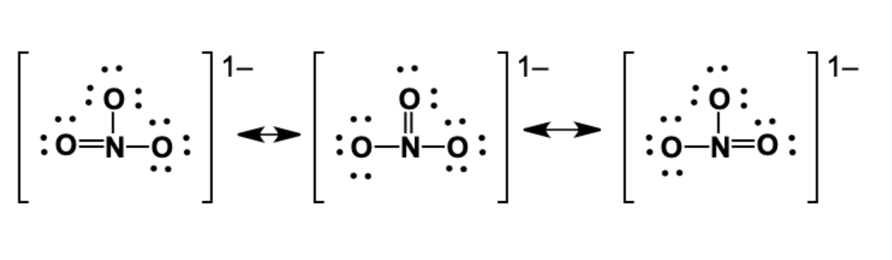


Exercise 1.6.6.3.1

Draw the Lewis (electron dot) structure of the nitrate ion (NO_3^{1-}).

Answer

Any of the three resonance structures shown below would work.



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1.6.7: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:

Figure 1.6.7.1: Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.

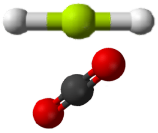


Figure 1.6.7.2: Beryllium hydride and carbon dioxide models. (CK12 Licence)

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

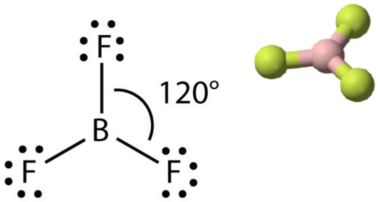


Figure 1.6.7.3: Boron trifluoride bonding. (CK12 Licence)

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

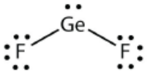


Figure 1.6.7.4: Germanium difluoride bonding.

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 1.6.7.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

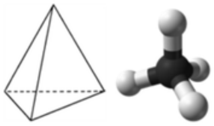


Figure 1.6.7.5: Tetrahedral structure of methane. (CK12 Licence)

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

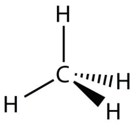


Figure 1.6.7.6: Methane bonding. (CK12 Licence)

NH_3 is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.

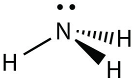


Figure 1.6.7.7: Ammonia bonding. (CK12 Licence)

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups, but only two of them are bonded to surrounding atoms.

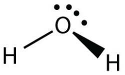


Figure 1.6.7.8: Water bonding.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom, is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 1.6.7.9

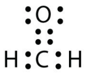


Figure 1.6.7.9: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

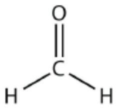


Figure 1.6.7.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table 1.6.7.1summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table 1.6.7.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

✓ Example 1.6.7.1

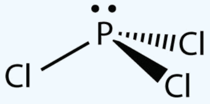
What is the approximate shape of each molecule?

a. PCl3
b. NOF

Solution

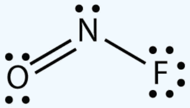
The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

? Exercise 1.6.7.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

? Exercise 1.6.7.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. (Hint: hydrogen is a terminal atom.)

Answer

Trigonal planar about both central C atoms.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

1.6.7.1: Practice Shape

Exercise 1.6.7.1.1

Give the electron geometry and the molecule geometry of PCl_3. (The Lewis structure was on a previous practice page.)

Answer

Electron geometry is tetrahedral (109.5° angle) because there are 4 electron regions around the central P atom.

Molecular geometry is pyramidal (also called trigonal pyramidal) because there are 3 atoms and 1 lone pair around the central P atom.

Exercise 1.6.7.1.1

Give the electron geometry and the molecule geometry of H_2O.

Answer

Electron geometry is tetrahedral (109.5° angle) because there are 4 electron regions around the central O atom.

Molecular geometry is bent because there are 2 atoms and 2 lone pairs around the central O atom.

Exercise 1.6.7.1.1

Give the electron geometry and the molecule geometry of the nitrate ion (NO_3^{1-}). (The Lewis structure was on a previous practice page.)

Answer

Electron geometry is trigonal planar (120° angle) because there are 3 electron regions around the central N atom.

Molecular geometry is trigonal planar because all 3 electron regions around the central N atom are atoms (no lone pairs).

1.6.8: Electronegativity and Polarity - Why Oil and Water Don’t Mix

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom, and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur), and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 1.6.8.1.

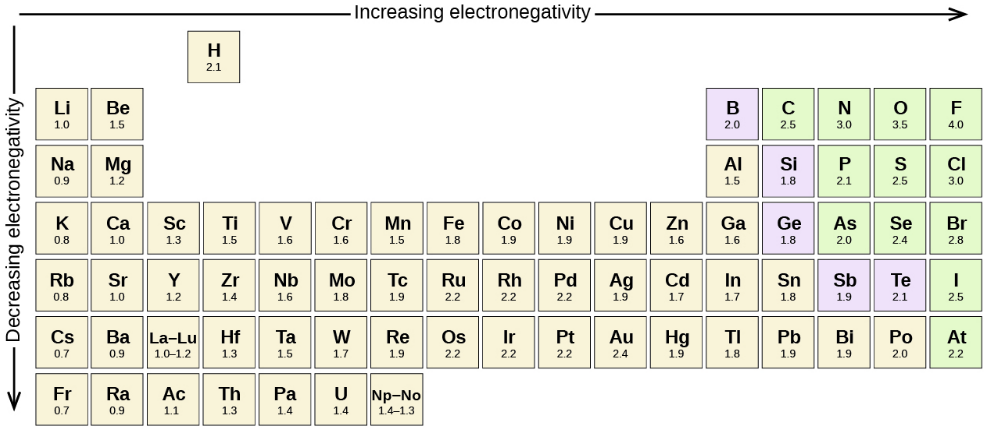


Figure 1.6.8.1: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

difference of the electronegativities of the two atoms involved in the covalent bond

Electronegativity Difference	Bond Type
0–0.4	pure covalent
0.5–2.0	polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

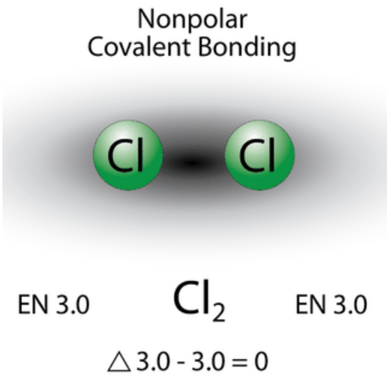


Figure 1.6.8.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathrm{Cl_2}$ molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta\mathrm{EN} = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

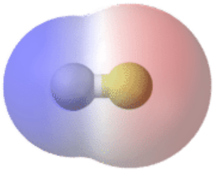


Figure 1.6.8.3: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

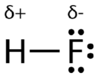


Figure 1.6.8.4: Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

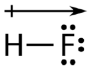


Figure 1.6.8.5: Use of crossed arrow to indicate polarity.

Electronegativity differences in bonding using Pauling scale. Differences in electronegativity classify bonds as covalent, polar covalent, or ionic.

✓ Example 1.6.8.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 1.6.8.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is $2.5 - 2.1 = 0.4$. Thus we predict that this bond will be nonpolar covalent.
b. For the O–H bond, the difference in electronegativities is $3.5 - 2.1 = 1.4$, so we predict that this bond will be polar covalent.

? Exercise 1.6.8.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is generally useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical—bonded to the same element with no unshared pairs of electrons. **Polar molecules** are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well, as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming that you do, you can look at the structure of each one and decide if it is polar or not, *whether or not you know the individual atom's electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.

Figure 1.6.8.6: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide (CO_2) is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of CO_2 is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

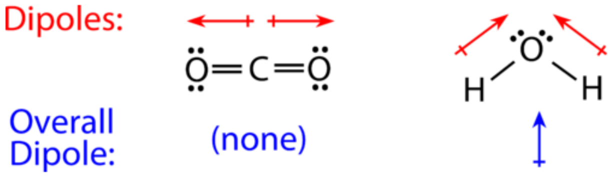


Figure 1.6.8.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as CH_4 is nonpolar. However, if one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule (BF_3) may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule (NH_3) is polar.

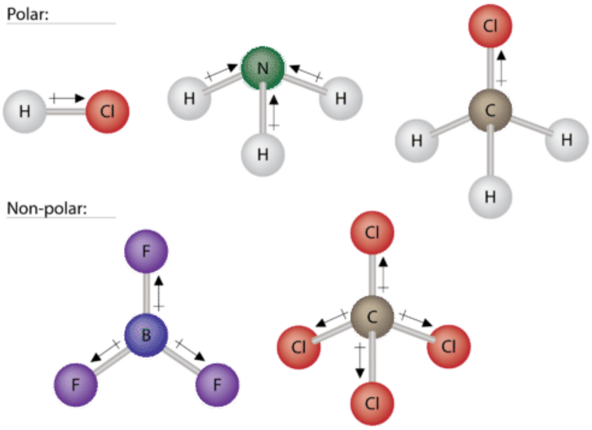


Figure 1.6.8.8: Some examples of polar and nonpolar molecules based on molecular geometry.

To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment do not cancel.

Steps to Identify Polar Molecules

1. Draw the Lewis structure.
2. Figure out the geometry (using VSEPR theory).
3. Visualize or draw the geometry.
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it).
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure 1.6.8.14). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

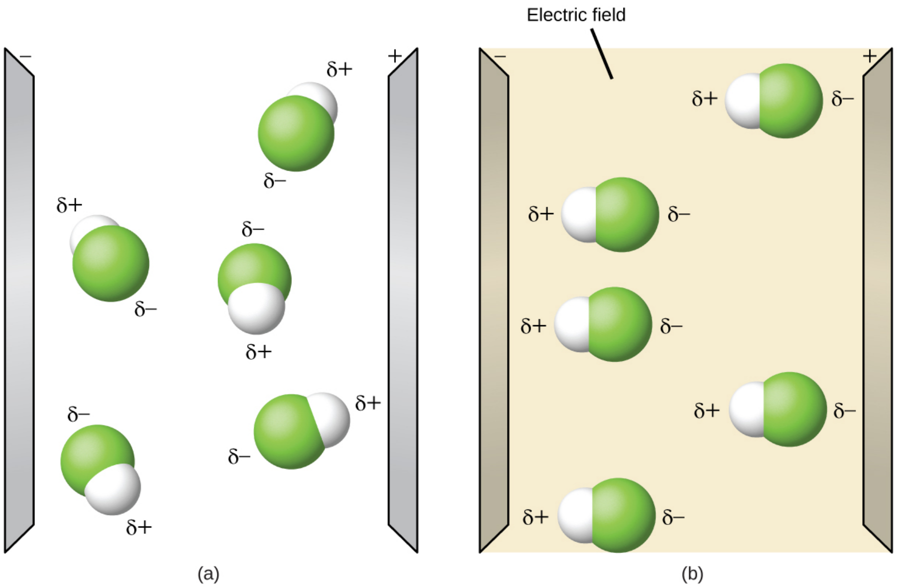


Figure 1.6.8.9: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

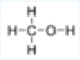
While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

✓ Example 1.6.8.2:

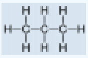
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

? Exercise 1.6.8.2

Label each of the following as polar or nonpolar.

a. SO_3
b. NH_3

Answer a

nonpolar

Answer b

polar

Contributions & Attributions

- StackExchange (thomij).

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1.6.8.1: Practice Polarity

Exercise 1.6.8.1.1

Is the PCl_3 molecule non-polar or polar? Check the electronegativity table in the text to see if the bonds are polar. Then see if the shape is symmetrical. Then decide.

Answer

The bonds are polar. Cl is more electronegative than P. Cl atoms are partially negative and P is partially positive.

The molecule is not entirely symmetrical, because the lone pair is causing the 3 Cl atoms to be slightly bent away from it, causing a pull of electrons toward the opposite side. The side of the molecule with the lone pair is slightly positive, the side with the chloringes is slightly negative. (See drawing of NH_3 molecule in text.)

It is polar.

Exercise 1.6.8.1.1

Is the CH_2O molecule non-polar or polar? Check the electronegativity table in the text to see if the bonds are polar. Then see if the shape is symmetrical. Then decide. (C and O have double bond. H atoms are attached to C with single bonds.)

Answer

The bonds between C and H are considered non-polar because the electronegativity difference is small (0.4). The bond between C and O is polar. The C is partially positive and the O is partially negative.

The molecule is not symmetrical because the oxygen is on one side of the carbon and the hydrogens are on the other side. The side of the molecule with the oxygen is slightly negative, and the side with the hydrogens is slightly positive.

It is polar.

Exercise 1.6.8.1.1

Is the H_2O molecule non-polar or polar? Check the electronegativity table in the text to see if the bonds are polar. Then see if the shape is symmetrical. Then decide.

Answer

The bonds between H and O are polar. The H are partially positive and the O is partially negative.

The molecule is not symmetrical because it is bent. The side with the lone pairs is slightly negative, and the side with the hydrogens is slightly positive.

It is polar.

SECTION OVERVIEW

1.7: Chemical Reactions and Quantities

1.7.1: The Mole

Learning Objectives

- Describe the unit *mole*.
- Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we do not typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. We need a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a specific number of things, just like a dozen means a certain number of things—twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A). *Avogadro's number* is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 1.7.1.1

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76\ \cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{\cancel{mol\ H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24}\ \cancel{molecules\ H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 1.7.1.1

How many molecules are present in 4.61×10^{-2} mol of O_2?

Answer

2.78×10^{22} molecules

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units*. The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022 × 10^{23} atoms:

$$12\text{ g C} = 6.022 \times 10^{23}\text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50 × 10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25}\ \cancel{atoms\,C} \times \frac{12.0000\,g\,C}{6.022 \times 10^{23}\ \cancel{atoms\,C}} = 299\,g\,C$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2. One molecule has a mass of 1.0079 + 1.0079 = 2.0158 u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022 × 10^{23} units: 6.022 × 10^{23} atoms of H, 6.022 × 10^{23} molecules of H_2 and H_2O, 6.022 × 10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

✓ Example 1.7.1.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula; but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

Per convention, the unit *grams per mole* is written as a fraction.

? Exercise 1.7.1.2

What is the molar mass of $AgNO_3$?

Answer

169.87 g/mol

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

✓ Example 1.7.1.3

What is the mass of 3.56 mol of $HgCl_2$? The molar mass of $HgCl_2$ is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56\ \cancel{mol\, HgCl_2} \times \frac{271.49\, g\, HgCl_2}{\cancel{mol\, HgCl_2}} = 967\, g\, HgCl_2$$

? Exercise 1.7.1.3

What is the mass of 33.7 mol of H_2O?

Answer

607 g

✓ Example 1.7.1.4

How many moles of H_2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles. The molar mass of water is

(1.0079 + 1.0079 + 15.999) = 18.015 g/mol.

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0\ \cancel{g\, H_2O} \times \frac{1\, mol\, H_2O}{18.015\ \cancel{g\, H_2O}} = 13.32\, mol\, H_2O$$

? Exercise 1.7.1.4

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer

0.363 mol

Other conversion factors can be combined with the definition of mole—density, for example.

✓ Example 1.7.1.5

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100\ \cancel{ml}\ ethanol \times \frac{0.789\, g}{\cancel{ml}} \times \frac{1\, mol}{46.08\ \cancel{g}} = 1.71\, mol\, ethanol$$

? Exercise 1.7.1.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer

0.201 mol

Summary

The mole is a key unit in chemistry. The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

1.7.1.1: Chemical Formulas as Conversion Factors

Learning Objectives

- Use chemical formulas as conversion factors.

Figure 1.7.1.1.1 shows that we need 2 hydrogen atoms and 1 oxygen atom to make one water molecule. If we want to make two water molecules, we will need 4 hydrogen atoms and 2 oxygen atoms. If we want to make five molecules of water, we need 10 hydrogen atoms and 5 oxygen atoms. The ratio of atoms we will need to make any number of water molecules is the same: 2 hydrogen atoms to 1 oxygen atom.

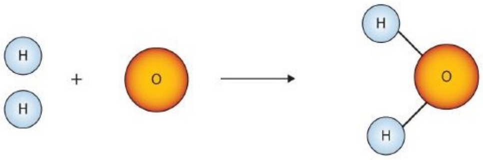


Figure 1.7.1.1.1 Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

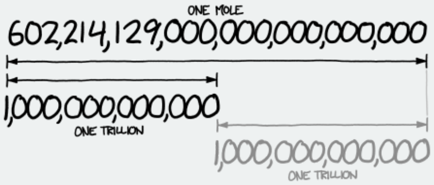
Using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of moles of molecules to the number of moles of atoms. For example, in 1 mol of water (H_2O) we can construct the relationships given in (Table 1.7.1.1.1).

Table 1.7.1.1.1 : Molecular Relationships for Water

1 Molecule of H_2O Has	1 Mol of H_2O Has	Molecular Relationships
2 H atoms	2 mol of H atoms	$\frac{2\text{ mol H atoms}}{1\text{ mol H}_2\text{O molecules}}$ or $\frac{1\text{ mol H}_2\text{O molecules}}{2\text{ mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\text{ mol O atoms}}{1\text{ mol H}_2\text{O molecules}}$ or $\frac{1\text{ mol H}_2\text{O molecules}}{1\text{ mol O atoms}}$

The Mole is big

A mole represents a very large number! The number 602,214,129,000,000,000,000,000 looks about twice as long as a trillion, which means it's about a trillion trillion.



(CC BY-SA NC; what if? [what-if.xkcd.com]).

A trillion trillion kilograms is how much a planet weighs. If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Table 1.7.1.1.2 : Molecular and Mass Relationships for Ethanol

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular and Mass Relationships
2 C atoms	2 mol of C atoms	$\frac{2\text{ mol C atoms}}{1\text{ mol C}_2\text{H}_6\text{O molecules}}$ or $\frac{1\text{ mol C}_2\text{H}_6\text{O molecules}}{2\text{ mol C atoms}}$
6 H atoms	6 mol of H atoms	$\frac{6\text{ mol H atoms}}{1\text{ mol C}_2\text{H}_6\text{O molecules}}$ or $\frac{1\text{ mol C}_2\text{H}_6\text{O molecules}}{6\text{ mol H atoms}}$

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular and Mass Relationships
1 O atom	1 mol of O atoms	$\frac{1 \text{ mol O atoms}}{1 \text{ mol } C_2H_6O \text{ molecules}}$ or $\frac{1 \text{ mol } C_2H_6O \text{ molecules}}{1 \text{ mol O atoms}}$
2 (12.01 amu) C 24.02 amu C	2 (12.01 g) C 24.02 g C	$\frac{24.02 \text{ g C}}{1 \text{ mol } C_2H_6O \text{ molecules}}$ or $\frac{1 \text{ mol } C_2H_6O \text{ molecules}}{24.02 \text{ g C}}$
6 (1.008 amu) H 6.048 amu H	6 (1.008 g) H 6.048 g H	$\frac{6.048 \text{ g H}}{1 \text{ mol } C_2H_6O \text{ molecules}}$ or $\frac{1 \text{ mol } C_2H_6O \text{ molecules}}{6.048 \text{ g H}}$
1 (16.00 amu) O 16.00 amu O	1 (16.00 g) O 16.00 g O	$\frac{16.00 \text{ g O}}{1 \text{ mol } C_2H_6O \text{ molecules}}$ or $\frac{1 \text{ mol } C_2H_6O \text{ molecules}}{16.00 \text{ g O}}$

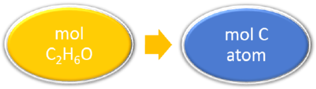
The following example illustrates how we can use the relationships in Table 1.7.1.1.2as conversion factors.

Example 1.7.1.1.1: Ethanol

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?

Solution

Solutions to Example 6.5.1

Steps for Problem Solving	If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?
Identify the "given" information and what the problem is asking you to "find."	Given: 2.5 mol C_2H_6O Find: mol C atoms
List other known quantities.	1 mol C_2H_6O = 2 mol C
Prepare a concept map and use the proper conversion factor.	 $\frac{2 \, mol \, C}{1 \, mol \, C_2H_6O}$
Cancel units and calculate.	Note how the unit *mol C_2H_6O molecules* cancels algebraically. $2.5 \text{ mol } \cancel{C_2H_6O \text{ molecules}} \times \frac{2 \text{ mol C atoms}}{1 \text{ mol } \cancel{C_2H_6O \text{ molecules}}} = 5.0 \text{ mol C atoms}$
Think about your result.	There are twice as many C atoms in one C_2H_6O molecule, so the final amount should be double.

Exercise 1.7.1.1.1

If a sample contains 6.75 mol of Na_2SO_4, how many moles of sodium atoms, sulfur atoms, and oxygen atoms does it have?

Answer

13.5 mol Na atoms, 6.75 mol S atoms, and 27.0 mol O atoms

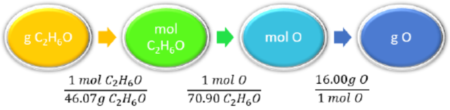
The fact that 1 mol equals 6.022×10^{23} items can also be used as a conversion factor.

Example 1.7.1.1.2: Oxygen Mass

Determine the mass of Oxygen in 75.0g of C_2H_6O.

Solution

Solutions to Example 6.5.2

Steps for Problem Solving	Determine the mass of Oxygen in 75.0g of C_2H_6O
Identify the "given" information and what the problem is asking you to "find."	Given: 75.0g C_2H_6O Find: g O
List other known quantities.	1 mol O = 16.0g O 1 mol C_2H_6O = 1 mol O 1 mol C_2H_6O = 46.07g C_2H_6O
Prepare a concept map and use the proper conversion factor.	 The conversion factors are 1 mol C2H6O over 46.07 g C2H6O, 1 mol O over 1 mol C2H6O, and 16.00 g O over 1 mole O.
Cancel units and calculate.	$75.0\ \text{g}\ \cancel{C_2H_6O} \times \frac{1\ \text{mol}\ \cancel{C_2H_6O}}{46.07\ \text{g}\ \cancel{C_2H_6O}} \times \frac{1\ \text{mol}\ \cancel{O}}{1\ \text{mol}\ \cancel{C_2H_6O}} \times \frac{16.00\ \text{g O}}{1\ \text{mol}\ \cancel{O}} = 26.0$
Think about your result.	

? Exercise 1.7.1.1.2

a. How many molecules are present in 16.02 mol of C_4H_{10}? How many C atoms are in 16.02 mol?
b. How many moles of each type of atom are in 2.58 mol of Na_2SO_4?

Answer a:

9.647×10^{24} C_4H_{10} molecules and 3.859×10^{25} C atoms

Answer b:

5.16 mol Na atoms, 2.58 mol S atoms, and 10.3 mol O atoms

Summary

In any given formula, the ratio of the number of moles of molecules (or formula units) to the number of moles of atoms can be used as a conversion factor.

1.7.1.2: Practice Mole Calculations

Molar Mass, Gram to Mole, Mole to Gram

Exercise 1.7.1.2.1

How many moles of $Ca_3(PO_4)_2$ are in 57.82 grams?

Molar Mass

$$Ca_3(PO_4)_2 = \frac{310.18g}{mol}$$

Answer

0.1864 moles

How many grams of $Zn(OH)_2$ are in 3.67 moles?

Molar Mass

$$Zn(OH)_2 = \frac{99.41g}{mol}$$

Answer

365 grams

How many moles of $Mg(NO_3)_2$ are in 487.6 grams?

Molar Mass

$$Mg(NO_3)_2 = \frac{148.33g}{mol}$$

Answer

3.287 moles

Using Avogadro's Number

Exercise 1.7.1.2.1

How many molecules of H_2O are in 0.0643 grams of H_2O?

Answer

0.0643 grams of H_2O is equal to 0.00357 moles, which is 2.15 x 10^{21} molecules

How many moles of Au is equal to 4.77 x 10^{27} Au atoms?

Answer

7920 moles

Using Formula Subscripts

Exercise 1.7.1.2.1

How many moles of NH_3 would contain 8.18 moles of H atoms?

Answer

2.73 moles NH_3

How many Br atoms are in 0.453 moles of CBr_4?

Answer

1.09 x 10^{24} Br atoms

Mixed

Exercise 1.7.1.2.1

How many moles of $Ca(OH)_2$ are in 5.62 g of $Ca(OH)_2$?

Answer

0.0758 mol $Ca(OH)_2$

How many moles of O are in this amount?

Answer

0.152 mol O

How many individual atoms of O are in this amount?

Answer

9.13 x 10^{22} O atoms

Exercise 1.7.1.2.1

How many moles of $C_4H_8O_2$ can be made with 18.5 mol C?

Answer

4.63 moles of $C_4H_8O_2$

Exercise 1.7.1.2.1

How many grams of $C_4H_8O_2$ can be made with 2.86 x 10^{21} C atoms? (Hint: how many moles of C is this?)

Hint

0.00475 mol C

Answer

0.00119 mol $C_4H_8O_2$ which is 0.105 g $C_4H_8O_2$

1.7.2: Percent Composition

Nutrition Facts

Serv. Size 2 tbsp (32g)
Servings about 16

Calories 180
Fat Cal. 130

*Percent Daily Values (DV) are based on a 2,000 calorie diet.

Amount Per Serving	% DV*	Amount Per Serving	% DV*
Total Fat 15g	23%	**Total Carbohydrate** 8g	3%
Saturated Fat 2g	10%	Dietary Fiber 2g	8%
Trans Fat 0g		Sugars 3g	
Cholesterol 0mg	0%	**Protein** 7g	
Sodium 140mg	6%		

Vitamin A 0% • Vitamin C 0% • Calcium 2% • Iron 4%

INGREDIENTS: PEANUTS, SUGAR, HYDROGENATED VEGETABLE OIL (RAPESEED, COTTONSEED, OYBEAN), SALT, MOLASSES.

Figure 1.7.2.1 (Credit: Joy Sheng; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Is there anything healthy in this jar?

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter (shown above) reveals that one serving size is considered to be 32 g. The label also gives the masses of various types of compounds that are present in each serving. One serving contains 7 g of protein, 15 g of fat, and 3 g of sugar. By calculating the fraction of protein, fat, or sugar in one serving size of peanut butter and converting to percent values, we can determine the composition of peanut butter on a percent by mass basis.

Percent Composition

Chemists often need to know what elements are present in a compound and in what percentage. The **percent composition** is the percent by mass of each element in a compound. It is calculated in a similar way to what was just indicated for the peanut butter.

$$\% \text{ by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

Percent Composition from Mass Data

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

Example 1.7.2.1: Percent Composition from Mass

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a 20.00 g sample of the compound is decomposed, $16.07 \text{ g}$of zinc remains. Determine the percent composition of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass of compound $= 20.00 \text{ g}$
- Mass of Zn $= 16.07 \text{ g}$

Unknown

- Percent Zn $= ?\%$
- Percent O $= ?\%$

Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass.

Step 2: Calculate.

$$\text{Mass of oxygen} = 20.00 \text{ g} - 16.07 \text{ g} = 3.93 \text{ g O}$$

$$\% \text{ Zn} = \frac{16.07 \text{ g Zn}}{20.00 \text{ g}} \times 100\% = 80.35\% \text{ Zn}$$

$$\% \text{ O} = \frac{3.93 \text{ g O}}{20.00 \text{ g}} \times 100\% = 19.65\% \text{ O}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\% \text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 1.7.2.2: Percent Composition from Chemical Formula

Dichlorineheptoxide $\left(\mathrm{Cl_2O_7}\right)$ is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorineheptoxide.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass of Cl in 1 mol $\mathrm{Cl_2O_7} = 70.90 \text{ g}$
- Mass of O in 1 mol $\mathrm{Cl_2O_7} = 112.00 \text{ g}$
- Molar mass of $\mathrm{Cl_2O_7} = 182.90 \text{ g/mol}$

Unknown

- Percent $\mathrm{Cl} = ?\%$
- Percent $\mathrm{O} = ?\%$

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\%\mathrm{Cl} = \frac{70.90 \text{ g Cl}}{182.90 \text{ g}} \times 100\% = 38.76\% \text{ Cl}$$

$$\% \text{ O} = \frac{112.00 \text{ g O}}{182.90 \text{ g}} \times 100\% = 61.24\% \text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\% \text{ Cl}$ and $61.24\% \text{ O}$. Suppose that you need to know the masses of chlorine and oxygen present in a 12.50 g sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element:

$$12.50 \text{ g } \mathrm{Cl_2O_7} \times \frac{38.76 \text{ g Cl}}{100 \text{ g } \mathrm{Cl_2O_7}} = 4.845 \text{ g Cl}$$

$$12.50 \text{ g } \mathrm{Cl_2O_7} \times \frac{61.24 \text{ g O}}{100 \text{ g } \mathrm{Cl_2O_7}} = 7.655 \text{ g O}$$

The sum of the two masses is 12.50 g the mass of the sample size.

Science Friday: Stained Glass Conservation

Stained glass from the Middle Ages is often hundreds of years old. Unfortunately, many of these relics are in need of cleaning and maintenance. In this video by Science Friday, conservator Mary Higgins discusses the methods used to protect the stained glass.

Summary

- Processes are described for calculating the percent composition of a material based on mass or on chemical composition.

Review

1. What is the formula for calculating percent composition?
2. What information do you need to calculate percent composition by mass?
3. What do subscripts in a chemical formula tell you?

1.7.2.1: Determining Empirical Formulas

Figure 1.7.2.1.1 (Credit: Harriet Moore; Source: http://commons.wikimedia.org/wiki/File:M_Faraday_Lab_H_Moore.jpg(opens in new window); License: Public Domain)

What is occuring in this picture?

In the early days of chemistry, there were few tools available for the detailed study of compounds. Much of the information regarding the composition of compounds came from the elemental analysis of inorganic materials. The "new" field of organic chemistry (the study of carbon compounds) faced the challenge of not being able to characterize a compound completely. The relative amounts of elements could be determined, but many of these materials had carbon, hydrogen, oxygen, and possibly nitrogen in simple ratios. Chemists had no way to determine the exact amounts of these atoms that were contained in specific molecules.

Determining Empirical Formulas

An **empirical formula** is one that shows the lowest whole-number ratio of the elements in a compound. Because the structure of ionic compounds is an extended three-dimensional network of positive and negative ions, all formulas of ionic compounds are empirical. However, we can also consider the empirical formula of a molecular compound. Ethene is a small hydrocarbon compound with the formula $\mathrm{C_2H_4}$ (see figure below). While $\mathrm{C_2H_4}$ is its molecular formula and represents its true molecular structure, it has an empirical formula of $\mathrm{CH_2}$. The simplest ratio of carbon to hydrogen in ethene is 1:2. There are two ways to view that ratio. Considering one molecule of ethene, the ratio is 1 carbon atom for every 2 atoms of hydrogen. Considering one mole of ethene, the ratio is 1 mole of carbon for every 2 moles of hydrogen. So, the subscripts in a formula represent the mole ratio of the elements in that formula.

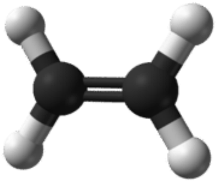


Figure 1.7.2.1.2: Ball-and-stick model of ethene, $\mathrm{C_2H_4}$. (Credit: Ben Mills (Wikimedia: Benjah-bmm27); Source: http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png(opens in new window); License: Public Domain)

In a procedure called **elemental analysis**, an unknown compound can be analyzed in the laboratory in order to determine the percentages of each element contained within it. These percentages can be transformed into the mole ratio of the elements, which leads to the empirical formula. The steps to be taken are outlined below.

1. Assume a $100\ \mathrm{g}$ sample of the compound, so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.
3. In order to find a whole-number ratio, divide the moles of each element by whichever of the moles from step 2 is the smallest.
4. If all the moles at this point are whole numbers (or very close), the empirical formula can be written with the moles as the subscript of each element.
5. In some cases, one or more of the moles calculated in step 3 will not be whole numbers. Multiply each of the moles by the smallest whole number that will convert each into a whole number. Write the empirical formula.

Example 1.7.2.1.1: Determining the Empirical Formula of a Compound

A compound of iron and oxygen is analyzed and found to contain 69.94% iron and 30.06% oxygen. Find the empirical formula of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\% \text{ of Fe} = 69.94\%$
- $\% \text{ of O} = 30.06\%$

Unknown

- Empirial formula = $Fe_?O_?$

Steps to follow are outlined in the text.

Step 2: Calculate.

1. Assume a 100 g sample.

$$69.94 \text{ g Fe}$$

$$30.06 \text{ g O}$$

2. Convert to moles.

$$69.94 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.252 \text{ mol Fe}$$

$$30.06 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.879 \text{ mol O}$$

3. Divide both moles by the smallest of the results.

$$\frac{1.252 \text{ mol Fe}}{1.252} = 1 \text{ mol Fe} \quad \frac{1.879 \text{ mol O}}{1.252} = 1.501 \text{ molO}$$

4/5. Since the moles of O is still not a whole number, both moles can be multiplied by 2, while rounding to a whole number.

$$1 \text{ mol Fe} \times 2 = 2 \text{ mol Fe} \quad 1.501 \text{ mol O} \times 2 = 3 \text{ mol O}$$

The empirical formula of the compound is Fe_2O_3.

Step 3: Think about your result.

The subscripts are whole numbers and represent the mole ratio of the elements in the compound. The compound is the ionic compound iron (III) oxide.

Summary

- A process is described for the calculation of the empirical formula for a compound based on the percent composition of that compound.

Review

1. What is an empirical formula?
2. What does an empirical formula tell you?
3. What does it not tell you?

1.7.2.2: Determining Molecular Formulas

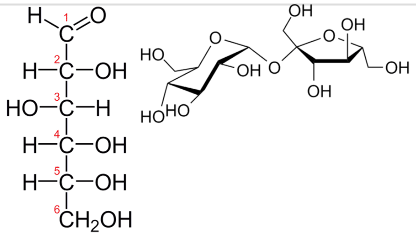


Figure 1.7.2.2.1 (Left: Credit: Ben Mills (Wikimedia: Benjah-bmm27) Source: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png(opens in new window); http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png(opens in new window) License: Public Domain) (Right: Credit: User:glycoform/Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27); Source: http://commons.wikimedia.org/wiki/File:Sucrose_3Dprojection.png(opens in new window); http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png(opens in new window); License: Public Domain)

How can you determine the differences between these two molecules?

Above we see two carbohydrates: glucose and sucrose. Sucrose is almost exactly twice the size of glucose, although their empirical formulas are very similar. Some people could distinguish them on the basis of taste, but it's not a good idea to go around tasting chemicals. The best way is to determine the molecular weights—this approach allows you to easily tell which compound is which.

Molecular Formulas

Molecular formulas give the kind and number of atoms of each element present in a molecular compound. In many cases, the molecular formula is the same as the empirical formula. The molecular formula of methane is $\mathrm{CH_4}$ and because it contains only one carbon atom, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole-number multiple of the empirical formula. Acetic acid is an organic acid that is the main component of vinegar. Its molecular formula is $\mathrm{C_2H_4O_2}$. Glucose is a simple sugar that cells use as a primary source of energy. Its molecular formula is $\mathrm{C_6H_{12}O_6}$. The structures of both molecules are shown in the figure below. They are very different compounds, yet both have the same empirical formula of $\mathrm{CH_2O}$.

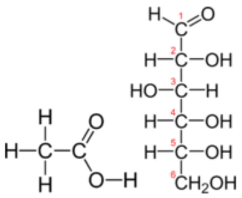


Figure 1.7.2.2.2: Acetic acid (left) has a molecular formula of $\mathrm{C_2H_4O_2}$, while glucose (right) has a molecular formula of $\mathrm{C_6H_{12}O_6}$. Both have the empirical formula $\mathrm{CH_2O}$. (Credit: (left) Ben Mills (Wikimedia: Benjah-bmm27); (right) Ben Mills (Wikimedia: Benjah-bmm27), User:Yikrazuul/Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27); Source: Acetic acid: http://commons.wikimedia.org/wiki/File:Acetic-acid-2D-flat.png(opens in new window); Glucose: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png(opens in new window); http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png(opens in new window); License: Public Domain)

Empirical formulas can be determined from the percent composition of a compound. In order to determine its molecular formula, it is necessary to know the molar mass of the compound. Chemists use an instrument called a mass spectrometer to determine the molar mass of compounds. In order to go from the empirical formula to the molecular formula, follow these steps:

1. Calculate the **empirical formula mass (EFM)**, which is simply the molar mass represented by the empirical formula.
2. Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.
3. Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.

Example 1.7.2.2.1: Determining the Molecular Formula of a Compound

The empirical formula of a compound of boron and hydrogen is $\mathrm{BH_3}$. Its molar mass is $27.7\ \mathrm{g/mol}$ Determine the molecular formula of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Empirical formula $= \mathrm{BH_3}$
- Molar mass $= 27.7\ \mathrm{g/mol}$

Unknown

- molecular formula = ?

Steps to follow are outlined in the text.

Step 2: Calculate.

$$\text{Empirical formula mass (EFM)} = 13.84\ \mathrm{g/mol}$$

$$\frac{\text{molar mass}}{\text{EFM}} = \frac{17.7}{13.84} = 2$$

$$\mathrm{BH_3} \times 2 = \mathrm{B_2H_6}$$

The molecular formula of the compound is $\mathrm{B_2H_6}$.

Step 3: Think about your result.

The molar mass of the molecular formula matches the molar mass of the compound.

Summary

- A procedure is described for the calculation of the exact molecular formula of a compound.

Review

1. What is the difference between an empirical formula and a molecular formula?
2. In addition to the elemental analysis, what do you need to know to calculate the molecular formula?
3. What does the empirical formula mass tell you?

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1.7.2.3: Practice Empirical and Molecular Formulas

Percent Composition Problems

Exercise 1.7.2.3.1

What is the percent composition of $Co(NO_3)_2$?

Hint: What is formula mass?

$\text{Co mass } = 1 \times 58.93 = 58.93\,\text{amu}$

$\text{N mass } = 2 \times 14.01 = 28.02\,\text{amu}$

$\text{O mass } = 6 \times 16.00 = 96.00\,\text{amu}$

Total = 182.95 amu

Answer

Co mass percent $= \dfrac{58.93 amu}{182.95 amu} \times 100$ % = 32.21 %

N mass percent $= \dfrac{28.02 amu}{182.95 amu} \times 100$ % = 15.32 %

O mass percent $= \dfrac{96.00 amu}{182.95 amu} \times 100$ % = 52.47 %

What is the percent composition of $Ni_2(SO_4)_3$?

Hint: What is formula mass?

$\text{Ni mass } = 2 \times 58.69 = 117.38\,\text{amu}$

$\text{S mass } = 3 \times 32.07 = 96.21\,\text{amu}$

$\text{O mass } = 12 \times 16.00 = 192.00\text{amu}$

Total = 405.59 amu

Answer

Ni mass percent $= \dfrac{117.38 amu}{405.59 amu} \times 100$ % = 28.94 %

S mass percent $= \dfrac{96.21 amu}{405.59 amu} \times 100$ % = 23.72 %

O mass percent $= \dfrac{192.00 amu}{405.59 amu} \times 100$ % = 47.34 %

You analyze a sample and find that it contains 22.46 g of Fe and 42.76 g Cl. What is the percent composition of this sample?

Hint: What is total mass?

Your sample is 22.46 g + 42.76 g = 65.22 g

Answer

Fe mass percent $= \dfrac{22.46 g}{65.22 g} \times 100$ % = 34.44 %

Cl mass percent $= \dfrac{42.76 g}{65.22 g} \times 100$ % = 65.56 %

Empirical Formula from Percent Composition

Exercise 1.7.2.3.1

What is the empirical formula of a compound that is 73.42 % Co and 26.58% O by mass?

Hint

Convert 73.42 g Co to moles Co (Co = 58.93 g/mol) and 26.58 g O to moles O (O = 16.00 g/mol).

Answer

At first, you get 1.246 mol Co and 1.661 mol O. Divide both of these by the smaller number (1.246)

This gives 1 mol Co and 1.33 mol O. Now multiply both of these by 3.

This gives 3 mol Co and 4 mol O.

Co_3O_4

What is the empirical formula of a compound that is 26.52 % Cr, 24.53 % S, and 48.95% O by mass?

Hint

Convert 26.52 g Cr to moles Cr (Cr = 52.00 g/mol), 24.53 g S to moles S (S = 32.07 g/mol), and 48.95 g O to moles O (O = 16.00 g/mol).

Answer

At first, you get 0.5100 mol Cr, 0.7649 mol S, and 3.059 mol O. Divide all of these by the smallest number (0.5100)

This gives 1 mol Cr, 1.5 mol S, and 6 mol O. Now multiply all of these by 2.

This gives 2 mol Cr, 3 mol S, and 12 mol O.

$Cr_2S_3O_{12}$ (this is actually the compound $Cr_2(SO_4)_3$)

Empirical and Molecular Formulas

Exercise 1.7.2.3.1

For each formula below, give the empirical formula. Sometimes the formula given is the same as the empirical, sometimes it is different.

a) $C_3H_6O_3$

b) N_2O_4

c) Mg_3N_2

d) $C_7H_{14}O_2$

e) P_2O_5

f) $C_4H_8N_2$

Answer a and b

CH_2O and NO_2

Answer c and d

Mg_3N_2 and $C_7H_{14}O_2$

Answer e and f

P_2O_5 and C_2H_4N

What is the molecular formula for a compound with empirical formula of CH_2O and molecular mass of 150.15 amu?

Answer

$C_5H_{10}O_5$ (Molecular mass is five times as big as empirical formula mass.)

What is the molecular formula for a compound with empirical formula of CH_2NO_2 and molecular mass of 180.12 amu?

Answer

$C_3H_6N_3O_6$ (Molecular mass is three times as big as empirical formula mass.)

What is the molecular formula for a compound with a molecular mass of 86.18 amu that is found to be 83.62 % C and 16.38 % H by mass? (Must find empirical formula first using percent composition.)

Empirical Formula

C_3H_7

At first, you get 6.963 mol C and 16.22 mol H. Divide both of these by the smallest number (6.963)

This gives 1 mol C and 2.33 mol H. Now multiply both of these by 3.

This gives 3 mol C and 7 mol H.

Molecular Formula

C_6H_{14} (Molecular mass is two times as big as empirical formula mass.)

1.7.3: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products; and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

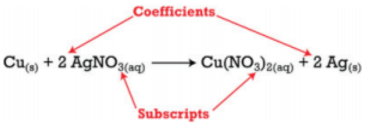
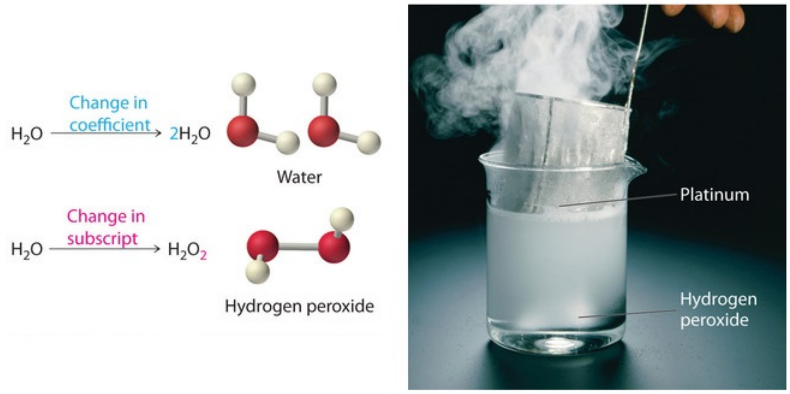


Figure 1.7.3.1: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 1.7.3.1



Original molecule H_2O: if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make H_2O_2, that's hydrogen peroxide.

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

Example 1.7.3.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).

$$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)}$$

Solution

Solutions to Example 1.7.3.1

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 $\mathrm{CO_2}$ molecules, each of which contains 1 carbon atom, on the right side: $\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow \underline{7}CO_2(g) + H_2O(g)}$ - 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 $\mathrm{H_2O}$ molecules, each of which contains 2 hydrogen atoms, on the right side: $\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow 7\,CO_2(g) + \underline{8}H_2O(g)}$ - 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $\mathrm{C_7H_{16}(l) + \underline{11}O_2(g) \rightarrow 7\,CO_2(g) + 8\,H_2O(g)}$ - 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

✓ Example 1.7.3.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Solutions to Example 7.4.2

Steps	**Example**
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $$C_8H_{18}(l) + \underline{\frac{25}{2}}O_2(g) \rightarrow 8\,CO_2(g) + 9\,H_2O(g)$$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $$\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$$
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Example 1.7.3.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Solutions to Example 7.4.3

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $\mathrm{Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)}$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $\mathrm{Pb(NO_3)_2(aq) + \underline{2}NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)}$ - 1 Pb atom on both reactant and product sides - 2 Na atoms on reactant side, 1 Na atom on product side - 2 Cl atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow \underline{2}NaNO_3(aq) + PbCl_2(s)}$ - 1 Pb atom on both reactant and product sides - 2 Na atoms on both reactant and product sides - 2 Cl atoms on both reactant and product sides - 2 NO_3^- atoms on both reactant and product sides
4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow 2\,NaNO_3(aq) + PbCl_2(s)}$ - 1 Pb atom on both reactant and product sides - 2 Na atoms on both reactant and product sides - 2 Cl atoms on both reactant and product sides - 2 NO_3^- atoms on both reactant and product sides

Exercise 1.7.3.1

Is each chemical equation balanced?

a. $\mathrm{2\,Hg(\ell)^{+}O_2(g) \rightarrow Hg_2O_2(s)}$
b. $\mathrm{C_2H_4(g) + 2\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)}$
c. $\mathrm{Mg(NO_3)_2(s) + 2\,Li(s) \rightarrow Mg(s)^{+}2\,LiNO_3(s)}$

Answer a

yes

Answer b

no

Answer c

yes

Exercise 1.7.3.2

Balance the following chemical equations.

a. $N_2(g) + O_2(g) \rightarrow NO_2(g)$
b. $Pb(NO_3)_2(aq) + FeCl_3(aq) \rightarrow Fe(NO_3)_3(aq) + PbCl_2(s)$
c. $C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Answer a

N_2 (g) + $2O_2$ (g) → $2NO_2$ (g)

Answer b

$3Pb(NO_3)_2$(aq) + $2FeCl_3$(aq) → $2Fe(NO_3)_3$(aq) + $3PbCl_2$(s)

Answer c

$2C_6H_{14}$(l) + $19O_2$(g)→ $12CO_2$(g) + $14H_2O$(g)

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- **Chemical reaction -** The process in which one or more substances are changed into one or more new substances.
- **Reactants -** The starting materials in a reaction.
- **Products -** Materials present at the end of a reaction.
- **Balanced chemical equation -** A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts -** Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient -** A small whole number that appears in front of a formula in a balanced chemical equation.

1.7.3.1: Practice Balancing Reactions

Balance Chemical Reactions

Exercise 1.7.3.1.1

For each of the following, adjust the coefficients so the reaction is balanced.

$NO_{2\,(g)} + H_2O_{\,(l)} \rightarrow HNO_{3\,(aq)} + NO_{\,(g)}$

Answer

$3\ NO_{2\,(g)} + H_2O_{\,(l)} \rightarrow 2\ HNO_{3\,(aq)} + NO_{\,(g)}$

$N_{2\,(g)} + O_{2\,(g)} \rightarrow N_2O_{5\,(g)}$

Answer

$2\ N_{2\,(g)} + 5\ O_{2\,(g)} \rightarrow 2\ N_2O_{5\,(g)}$

$C_5H_{10}O_{\,(l)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{\,(g)}$

Answer

$C_5H_{10}O_{\,(l)} + 7\ O_{2\,(g)} \rightarrow 5\ CO_{2\,(g)} + 5\ H_2O_{\,(g)}$

$ZnCl_{2\,(aq)} + AgNO_{3\,(aq)} \rightarrow Zn(NO_3)_{2\,(aq)} + AgCl_{\,(s)}$

Answer

$ZnCl_{2\,(aq)} + 2\ AgNO_{3\,(aq)} \rightarrow Zn(NO_3)_{2\,(aq)} + 2\ AgCl_{\,(s)}$

$BaCl_{2\,(aq)} + Na_3PO_{4\,(aq)} \rightarrow Ba_3(PO_4)_{2\,(s)} + NaCl_{\,(aq)}$

Answer

$3\ BaCl_{2\,(aq)} + 2\ Na_3PO_{4\,(aq)} \rightarrow Ba_3(PO_4)_{2\,(s)} + 6\ NaCl_{\,(aq)}$

$C_6H_{10\,(l)} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{\,(g)}$

Answer

$2\ C_6H_{10\,(l)} + 17\ O_{2\,(g)} \rightarrow 12\ CO_{2\,(g)} + 10\ H_2O_{\,(g)}$

$Mg_{\,(s)} + HCl_{\,(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$

Answer

$Mg_{\,(s)} + 2\ HCl_{\,(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$

1.7.4: Types of Chemical Reactions - Single and Double Replacement Reactions

Learning Objectives

- Recognize chemical reactions as single-replacement reactions and double-replacement reactions.
- Use the periodic table, an activity series, or solubility rules to predict whether single-replacement reactions or double-replacement reactions will occur.

Up until now, we have presented chemical reactions as a topic, but we have not discussed how the products of a chemical reaction can be predicted. Here we will begin our study of certain types of chemical reactions that allow us to predict what the products of the reaction will be.

A **single-replacement reaction** is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products. Presented below:

$$2\,\mathrm{HCl(aq)} + \mathrm{Zn(s)} \rightarrow \mathrm{ZnCl_2(aq)} + \mathrm{H_2(g)}$$

is an example of a single-replacement reaction. The hydrogen atoms in HCl are replaced by Zn atoms, and in the process a new element—hydrogen—is formed. Another example of a single-replacement reaction is

$$2\,\mathrm{NaCl(aq)} + \mathrm{F_2(g)} \rightarrow 2\,\mathrm{NaF(s)} + \mathrm{Cl_2(g)}$$

Here the negatively charged ion changes from chloride to fluoride. A typical characteristic of a single-replacement reaction is that there is one element as a reactant and another element as a product.

Not all proposed single-replacement reactions will occur between two given reactants. This is most easily demonstrated with fluorine, chlorine, bromine, and iodine. Collectively, these elements are called the *halogens* and are in the next-to-last column on the periodic table (Figure $1.7.4.1$). The elements on top of the column will replace the elements below them on the periodic table, but not the other way around. Thus, the reaction represented by

$$\mathrm{CaI_2(s)} + \mathrm{Cl_2(g)} \rightarrow \mathrm{CaCl_2(s)} + \mathrm{I_2(s)}$$

will occur; but the reaction

$$\mathrm{CaF_2(s)} + \mathrm{Br_2(\ell)} \rightarrow \mathrm{CaBr_2(s)} + \mathrm{F_2(g)}$$

will not, because bromine is below fluorine on the periodic table. This is just one of many ways the periodic table helps us to understand chemistry.

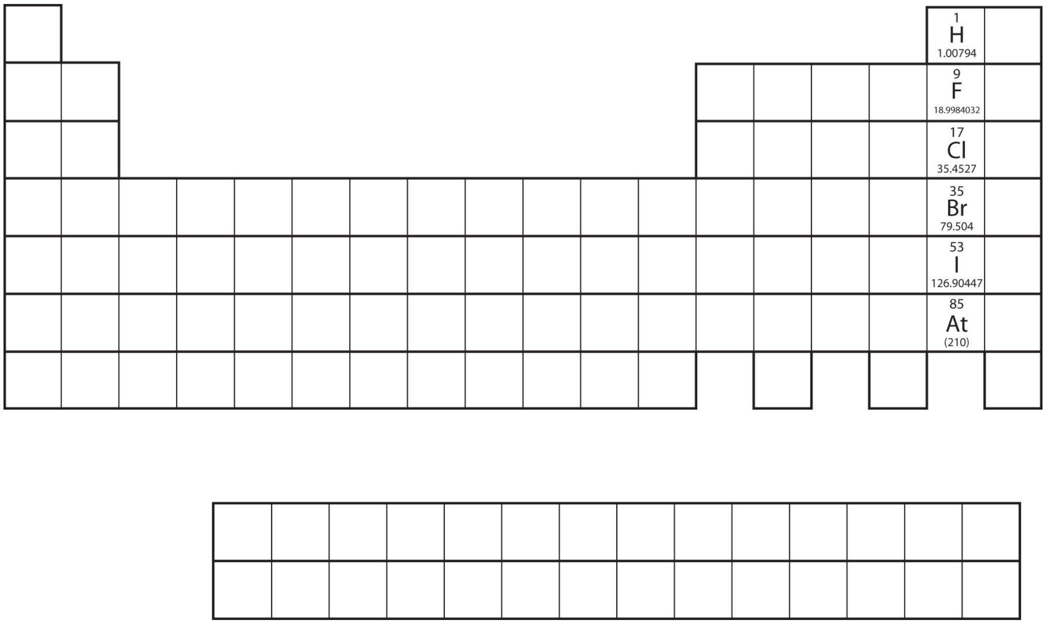


Figure 1.7.4.1: Halogens on the Periodic Table. The halogens are the elements in the next-to-last column on the periodic table.

Example 1.7.4.1

Will a single-replacement reaction occur? If so, identify the products.

a. $MgCl_2 + I_2 \rightarrow$?
b. $CaBr_2 + F_2 \rightarrow$?

Solution

a. Because iodine is below chlorine on the periodic table, a single-replacement reaction will not occur.
b. Because fluorine is above bromine on the periodic table, a single-replacement reaction will occur, and the products of the reaction will be CaF_2 and Br_2.

Exercise 1.7.4.1

Will a single-replacement reaction occur? If so, identify the products.

$$\mathrm{FeI_2 + Cl_2 \rightarrow}$$

Answer

Yes; $FeCl_2$ and I_2

Chemical reactivity trends are easy to predict when replacing anions in simple ionic compounds—simply use their relative positions on the periodic table. However, when replacing the cations, the trends are not as straightforward. This is partly because there are so many elements that can form cations; an element in one column on the periodic table may replace another element nearby, or it may not. A list called the **activity series** does the same thing the periodic table does for halogens: it lists the elements that will replace elements below them in single-replacement reactions. A simple activity series is shown below.

Activity Series for Cation Replacement in Single-Replacement Reactions

- Li
- K
- Ba
- Sr
- Ca
- Na
- Mg
- Al
- Mn
- Zn
- Cr
- Fe
- Ni
- Sn
- Pb
- H_2
- Cu
- Hg
- Ag
- Pd
- Pt
- Au

Using the activity series is similar to using the positions of the halogens on the periodic table. An element on top will replace an element below it in compounds undergoing a single-replacement reaction. Elements will not replace elements above them in compounds.

Example 1.7.4.2

Use the activity series to predict the products, if any, of each equation.

1. $FeCl_2$ + Zn → ?
2. HNO_3 + Au → ?

Solution

1. Because zinc is above iron in the activity series, it will replace iron in the compound. The products of this single-replacement reaction are $ZnCl_2$ and Fe.
2. Gold is below hydrogen in the activity series. As such, it will not replace hydrogen in a compound with the nitrate ion. No reaction is predicted.

Exercise 1.7.4.2

Use the activity series to predict the products, if any, of this equation.

$$\mathrm{AlPO_4 + Mg \rightarrow}$$

Answer

$Mg_3(PO_4)_2$ and Al

A **double-replacement reaction** occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is

$$\mathrm{CuCl_2(aq) + 2\,AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,AgCl(s)}$$

There are two equivalent ways of considering a double-replacement equation: either the cations are swapped, or the anions are swapped. (You cannot swap both; you would end up with the same substances you started with.) Either perspective should allow you to predict the proper products, as long as you pair a cation with an anion, and not a cation with a cation or an anion with an anion.

Example 1.7.4.3

Predict the products of this double-replacement equation:

$$\mathrm{BaCl_2 + Na_2SO_4 \rightarrow}$$

Solution

Thinking about the reaction as either switching the cations or switching the anions, we would expect the products to be $BaSO_4$ and NaCl.

Exercise 1.7.4.3

Predict the products of this double-replacement equation:

$$\mathrm{KBr + AgNO_3 \rightarrow}$$

Answer

KNO_3 and AgBr

Predicting whether a double-replacement reaction occurs is somewhat more difficult than predicting a single-replacement reaction. However, there is one type of double-replacement reaction that we can predict: the precipitation reaction. A **precipitation reaction** occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of solution as a solid **precipitate.** The formation of a solid precipitate is the driving force that makes the reaction proceed.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble, or insoluble). Table 1.7.4.1 lists some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules. If a compound is soluble, we use the (aq) label with it, indicating that it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

Table 1.7.4.1: Some Useful Solubility Rules (soluble)

These compounds generally dissolve in water (are soluble):	Exceptions:
All compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+	None
All compounds of NO_3^- and $C_2H_3O_2^-$	None
Compounds of Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds of SO_4^{2}	Hg_2^{2+}, Pb^{2+}, Sr^{2+}, Ba^{2+}

Table 1.7.4.2: Some Useful Solubility Rules (insoluble)

These compounds generally do not dissolve in water (are insoluble):	Exceptions:
Compounds of CO_3^{2-} and PO_4^{3-}	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+
Compounds of OH^-	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, Sr^{2+}, and Ba^{2+}

For example, consider the possible double-replacement reaction between Na_2SO_4 and $SrCl_2$. The solubility rules say that all ionic sodium compounds are soluble and all ionic chloride compounds are soluble, except for Ag^+, Hg_2^{2+}, and Pb^{2+}, which are not being considered here. Therefore, Na_2SO_4 and $SrCl_2$ are both soluble. The possible double-replacement reaction products are NaCl and $SrSO_4$. Are these soluble? NaCl is (by the same rule we just quoted), but what about $SrSO_4$? Compounds of the sulfate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:

$$\mathrm{Na_2SO_4(aq) + SrCl_2(aq) \rightarrow 2\,NaCl(aq) + SrSO_4(s)}$$

You would expect to see a visual change corresponding to $SrSO_4$ precipitating out of solution (Figure 1.7.4.2).

Figure 1.7.4.2: Double-Replacement Reactions. Some double-replacement reactions are obvious because you can see a solid precipitate coming out of solution. Source: Photo courtesy of Choij, http://commons.wikimedia.org/wiki/File:Copper_solution.jpg(opens in new window).

✓ Example 1.7.4.4

Will a double-replacement reaction occur? If so, identify the products.

1. $Ca(NO_3)_2$ + KBr → ?
2. NaOH + $FeCl_2$ → ?

Solution

1. According to the solubility rules, both $Ca(NO_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)—namely, $CaBr_2$ and KNO_3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and $FeCl_2$ are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $Fe(OH)_2$. NaCl is soluble, but, according to the solubility rules, $Fe(OH)_2$ is not. Therefore, a reaction would occur, and $Fe(OH)_2(s)$ would precipitate out of solution. The balanced chemical equation is

$$2\,\mathrm{NaOH(aq)} + \mathrm{FeCl_2(aq)} \rightarrow 2\,\mathrm{NaCl(aq)} + \mathrm{Fe(OH)_2(s)}$$

? Exercise 1.7.4.4

$$\mathrm{Sr(NO_3)_2} + \mathrm{KCl} \rightarrow$$

Answer

No reaction; all possible products are soluble.

Key Takeaways

- A single-replacement reaction replaces one element for another in a compound.
- The periodic table or an activity series can help predict whether single-replacement reactions occur.
- A double-replacement reaction exchanges the cations (or the anions) of two ionic compounds.
- A precipitation reaction is a double-replacement reaction in which one product is a solid precipitate.
- Solubility rules are used to predict whether some double-replacement reactions will occur.

1.7.4.1: Composition, Decomposition, and Combustion Reactions

Learning Objectives

- Recognize composition, decomposition, and combustion reactions.
- Predict the products of a combustion reaction.

Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easy to recognize.

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. In the reaction

$$2\,\mathrm{H_2(g)} + \mathrm{O_2(g)} \rightarrow 2\,\mathrm{H_2O}(\ell)$$

water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So this is a composition reaction.

A **decomposition reaction** starts from a single substance and produces more than one substance; that is, it decomposes. The key characteristics of a decomposition reaction are: one substance as a reactant and more than one substance as the products. For example, in the decomposition of sodium hydrogen carbonate (also known as sodium bicarbonate):

$$2\,\mathrm{NaHCO_3(s)} \rightarrow \mathrm{Na_2CO_3(s)} + \mathrm{CO_2(g)} + \mathrm{H_2O}(\ell)$$

sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict; however, they should be easy to recognize.

Example 1.7.4.1.1: Identifying Reactions

Identify each equation as a composition reaction, a decomposition reaction, or neither.

a. $\mathrm{Fe_2O_3} + 3\,\mathrm{SO_3} \rightarrow \mathrm{Fe_2(SO_4)_3}$
b. $\mathrm{NaCl} + \mathrm{AgNO_3} \rightarrow \mathrm{AgCl} + \mathrm{NaNO_3}$
c. $\mathrm{(NH_4)_2Cr_2O_7} \rightarrow \mathrm{Cr_2O_3} + 4\,\mathrm{H_2O} + \mathrm{N_2}$

Solution

a. In this equation, two substances combine to make a single substance. This is a composition reaction.
b. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
c. A single substance reacts to make multiple substances. This is a decomposition reaction.

Exercise 1.7.4.1.1

Identify the equation as a composition reaction, a decomposition reaction, or neither.

$$\mathrm{C_3H_8} \rightarrow \mathrm{C_3H_4} + 2\,\mathrm{H_2}$$

Answer

decomposition

A **combustion reaction** occurs when a reactant combines with oxygen, many times from the atmosphere, to produce oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N_2. Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen to produce CO_2 and H_2O. For example, the balanced chemical equation for the combustion of methane, CH_4, is as follows:

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O$$

Kerosene can be approximated with the formula $C_{12}H_{26}$, and its combustion equation is:

$$2\,C_{12}H_{26} + 37\,O_2 \rightarrow 24\,CO-2 + 26\,H_2O$$

Sometimes fuels contain oxygen atoms, which must be counted when balancing the chemical equation. One common fuel is ethanol, C_2H_5OH, whose combustion equation is:

$$C_2H_5OH + 3\,O_2 \rightarrow 2\,CO_2 + 3\,H_2O$$

If nitrogen is present in the original fuel, it is converted to N_2, not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $C_2H_2N_2O_4$, we have:

$$2\,C_2H_2N_2O_4 + O_2 \rightarrow 4\,CO_2 + 2\,H_2O + 2\,N_2$$

✓ Example 1.7.4.1.2: Combustion Reactions

Complete and balance each combustion equation.

1. the combustion of propane (C_3H_8)
2. the combustion of ammonia (NH_3)

Solution

1. The products of the reaction are CO2 and H2O, so our unbalanced equation is

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

Balancing (and you may have to go back and forth a few times to balance this), we get

$$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$$

2. The nitrogen atoms in ammonia will react to make N2, while the hydrogen atoms will react with O2 to make H2O:

$$NH_3 + O_2 \rightarrow N_2 + H_2O$$

To balance this equation without fractions (which is the convention), we get

$$4\,NH_3 + 3\,O_2 \rightarrow 2\,N_2 + 6\,H_2O$$

? Exercise 1.7.4.1.2

Complete and balance the combustion equation for cyclopropanol (\ce{C3H6O}\)).

Answer

$$C_3H_6O + 4\,O_2 \rightarrow 3\,CO_2 + 3\,H_2O$$

Figure 1.7.4.1.1 A 20 lb (9.1 kg) steel propane cylinder.

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.
- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products (although nitrogen atoms react to make N_2).

1.7.4.2: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate, a reddish precipitate of silver dichromate is produced.

$$\mathrm{AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + KNO_3(aq)} \tag{1.7.4.2.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$\mathrm{AC + BD \rightarrow \underbrace{AD}_{insoluble} + BC} \tag{1.7.4.2.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called **double-displacement reactions**. Precipitation reactions are used to isolate metals that have been extracted from their ores, and to recover precious metals for recycling.

Video: Mixing potassium dichromate and silver nitrate together to initiate a precipitation reaction (Equation 1.7.4.2.1).

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $NaCl$ solution is mixed with 500 mL of aqueous KBr solution, the final solution has a volume of 1.00 L and contains $\mathrm{Na^+(aq)}$, $\mathrm{Cl^-(aq)}$, $\mathrm{K^+(aq)}$, and $\mathrm{Br^-(aq)}$. As you will see in (Figure $1.7.4.2.1$), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

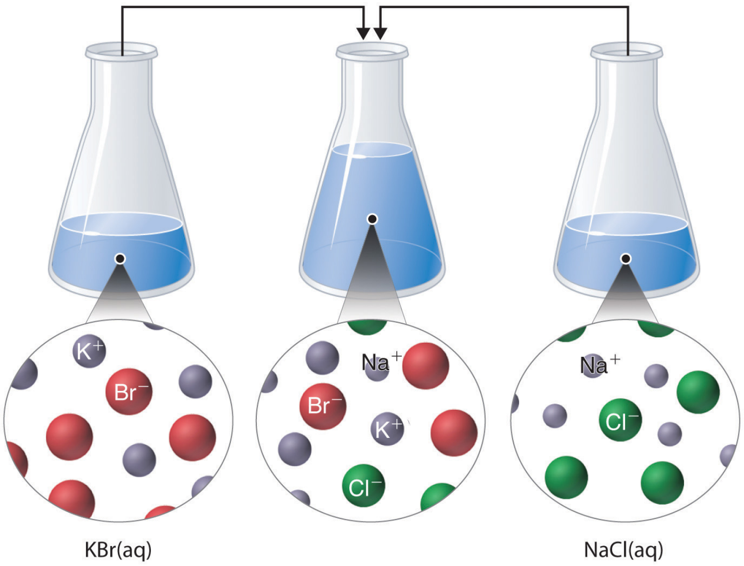


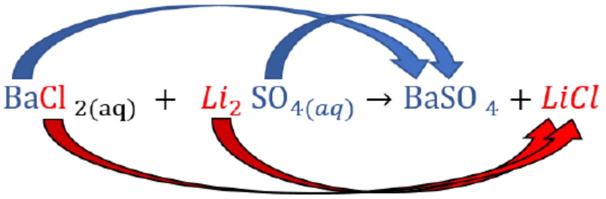
Figure 1.7.4.2.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

Change the partners of the anions and cations on the reactant side to form new compounds (products):



Chemical equation of the reactants barium chloride and lithium sulfate forming the products barium sulfate and lithium chloride.

Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains $\mathrm{Ba^{2+}}$, $\mathrm{Cl^-}$, $\mathrm{Li^+}$, and $\mathrm{SO_4^{2-}}$ ions. The only possible exchange reaction is to form LiCl and $\mathrm{BaSO_4}$.

Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$\mathrm{BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4 + LiCl}$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$\mathrm{BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + LiCl(aq)}$$

Table 7.5.1 from the previous section shows that LiCl is soluble in water, but $\mathrm{BaSO_4}$ is not soluble in water.

Balance the equation:

$$\mathrm{BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,LiCl(aq)}$$

Although soluble barium salts are toxic, $\mathrm{BaSO_4}$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a “barium milkshake” or a “barium enema”—a suspension of very fine $\mathrm{BaSO_4}$ particles in water.

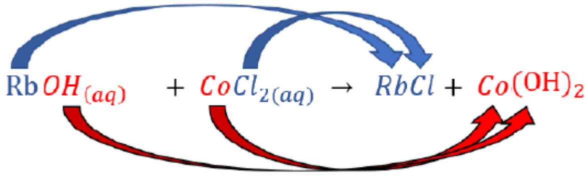
Figure 1.7.4.2.2: An x-ray of the digestive organs of a patient who has swallowed a “barium milkshake.” A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. (Public Domain; Glitzy queen00 via Wikipedia).

Example 1.7.4.2.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

Solution

Solutions to Example 7.6.1

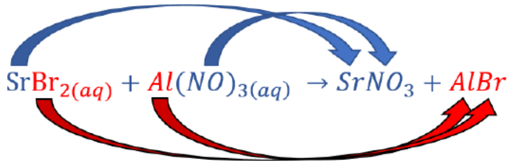
Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants rubidium hydroxide and cobalt(II) chloride forming the products rubidium chloride and cobalt hydroxide.
Correct the formulas of the products based on the charges of the ions.	$\text{RbOH}(\text{aq}) + \text{CoCl}_2(\text{aq}) \rightarrow \text{RbCl} + \text{Co}(\text{OH})_2$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\text{RbOH}(\text{aq}) + \text{CoCl}_2(\text{aq}) \rightarrow \text{RbCl}(\text{aq}) + \text{Co}(\text{OH})_2(\text{s})$
Balance the equation.	Coefficients already balanced. $\text{RbOH}(\text{aq}) + \text{CoCl}_2(\text{aq}) \rightarrow \text{RbCl}(\text{aq}) + \text{Co}(\text{OH})_2(\text{s})$

Example 1.7.4.2.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution

Solutions for Example 7.6.2

Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants strontium bromide and aluminum nitrate forming the products strontium nitrate and aluminum bromide.

Steps	Example
Correct the formulas of the products based on the charges of the ions.	$\mathrm{SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2 + AlBr_3}$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\mathrm{SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2(aq) + AlBr_3(aq)}$ According to Table 7.5.1 from the previous section, both $\mathrm{AlBr_3}$ (rule 4) and $\mathrm{Sr(NO_3)_2}$ (rule 2) are soluble.
If all possible products are soluble, then no net reaction will occur.	$\mathrm{SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow}$ **NO REACTION**

? Exercise 1.7.4.2.2

Using the information in Table 7.5.1 from the previous section, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a

$Fe(OH)_2$ precipitate is formed.

Answer b

$Hg_3(PO_4)_2$ precipitate is formed.

Answer c

No Reaction.

Answer d

$CaCO_3$ is precipitate formed.

Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributions & Attributions

- Modified by Joshua Halpern (Howard University)

1.7.4.3: Ionic Equations - A Closer Look

Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds—compounds between metals and nonmetals, or compounds that contained recognizable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolution in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process —for example, with NaCl:

$$\mathrm{NaCl(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+(aq)} + \mathrm{Cl^-(aq)}$$

When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 1.7.4.3.1). This process is called **dissociation;** we say that the ions *dissociate*.

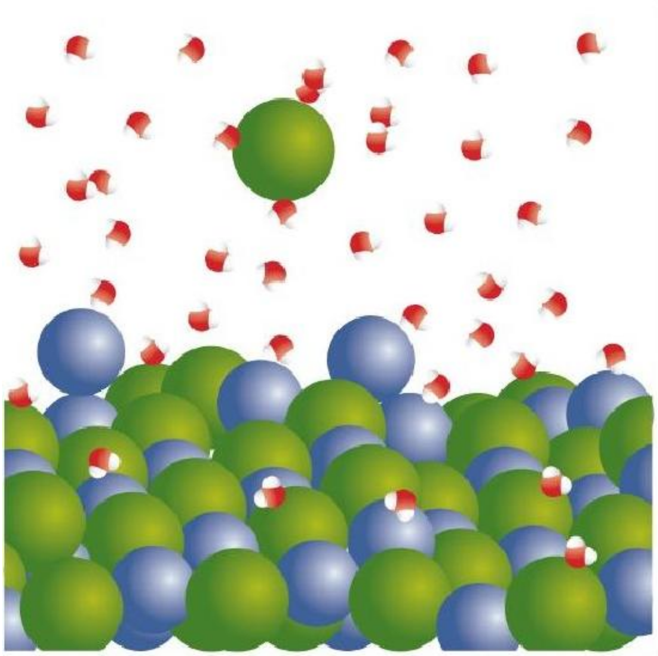


Figure 1.7.4.3.1 Ionic Solutions. When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. This behavior was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry. Keep in mind that when the ions separate, *all* the ions separate. Thus, when $CaCl_2$ dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from one another:

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + Cl^-(aq) + Cl^-(aq)$$

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + 2Cl^-(aq)$$

That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions), and they do not stick together to make Cl_2^- or Cl_2^{2-}. They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

✓ Example 1.7.4.3.1

Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solution

1. $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is $Na_2SO_4(s) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$

? Exercise 1.7.4.3.1

Write the chemical equation that represents the dissociation of $(NH_4)_2S$.

Answer

$(NH_4)_2S(s) \rightarrow 2NH_4^+(aq) + S^{2-}(aq)$

When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when $NaCl(aq)$ reacts with $AgNO_3(aq)$ in a double-replacement reaction to precipitate $AgCl(s)$ and form $NaNO_3(aq)$, the complete ionic equation includes NaCl, $AgNO_3$, and $NaNO_3$ written as separate ions:

$$\mathrm{Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)}$$

This is more representative of what is occurring in the solution.

✓ Example 1.7.4.3.1

Write the complete ionic equation for each chemical reaction.

1. $KBr(aq) + AgC_2H_3O_2(aq) \rightarrow KC_2H_3O_2(aq) + AgBr(s)$
2. $MgSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + BaSO_4(s)$

Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

1. The complete ionic equation is $K^+(aq) + Br^-(aq) + Ag^+(aq) + C_2H_3O_2^-(aq) \rightarrow K^+(aq) + C_2H_3O_2^-(aq) + AgBr(s)$
2. The complete ionic equation is $Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$

? Exercise 1.7.4.3.1

Write the complete ionic equation for

$$\mathrm{CaCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s)}$$

Answer

$Ca^{2+}(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + PbCl_2(s)$

You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in

$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

the $Ag^+(aq)$ and $Cl^-(aq)$ ions become $AgCl(s)$, but the $Na^+(aq)$ ions and the $NO_3^-(aq)$ ions stay as $Na^+(aq)$ ions and $NO_3^-(aq)$ ions. These two ions are examples of spectator ions—ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:

$$Na^+\cancel{(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-}(aq) \rightarrow AgCl(s) + \cancel{Na}^+(aq) + \cancel{NO}_3^-(aq)$$

What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:

$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

It is important to reiterate that the spectator ions are still present in solution, but they do not experience any net chemical change, so they are not written in a net ionic equation.

Example 1.7.4.3.1

Write the net ionic equation for each chemical reaction.

1. $K^+(aq) + Br^-(aq) + Ag^+(aq) + C_2H_3O_2^-(aq) \rightarrow K^+(aq) + C_2H_3O_2^-(aq) + AgBr(s)$
2. $Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$

Solution

1. In the first equation, the $K^+(aq)$ and $C_2H_3O_2^-(aq)$ ions are spectator ions, so they are canceled:

$$K^+\cancel{(aq)} + Br^-(aq) + Ag^+(aq) + \cancel{C_2H_3O_2^-(aq)} \rightarrow K^+(aq) + \cancel{C_2H_3O_2^-(aq)} + AgBr(s)$$

The net ionic equation is

$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s)$

2. In the second equation, the $Mg^{2+}(aq)$ and $NO_3^-(aq)$ ions are spectator ions, so they are canceled:

$$Mg^{2+}\cancel{(aq)} + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2\cancel{NO_3^-(aq)} \rightarrow Mg^{2+}(aq) + 2\cancel{NO_3^-(aq)} + BaSo_4(s)$$

The net ionic equation is

$SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$

Exercise 1.7.4.3.1

Write the net ionic equation for

$CaCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s)$

Answer

$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

Chemistry is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound that can be dissolved in a sample of water. For example, you can dissolve a maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider NaCl soluble but AgCl insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate ($CaCO_3$). However,

$CaCO_3$ has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water, $CaCO_3$ can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.

Figure 1.7.4.3.2 Tank water heater © Thinkstock. Most homes in the United States have a tank-type water heater like this one.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!

Figure 1.7.4.3.3: The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo. (Sonaal Bangera via unsplash)

Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

1.7.4.4: Oxidation-Reduction (Redox) Reactions

Learning Objectives

- To identify a chemical reaction as an oxidation-reduction reaction.

When zinc metal is submerged into a quantity of aqueous HCl, the following reaction occurs (Figure $1.7.4.4.1$):

$$\mathrm{Zn(s) + 2\,HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)} \tag{1.7.4.4.1}$$

This is one example of what is sometimes called a *single replacement reaction* because Zn replaces H in combination with Cl.

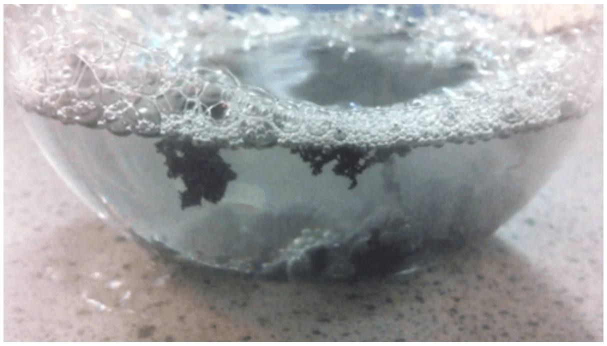


Figure $1.7.4.4.1$: Zinc Metal plus Hydrochloric Acid. It is fairly obvious that zinc metal reacts with aqueous hydrochloric acid! The bubbles are hydrogen gas (right side of Equation $1.7.4.4.1$).

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn(s) + 2\,H^+(aq) + 2\,Cl^-(aq) \rightarrow H_2(g) + Zn^{2+}(aq) + 2\,Cl^-(aq)}$$

Viewed this way, the net reaction seems to be a charge transfer between zinc and hydrogen atoms. (There is no net change experienced by the chloride ion.) In fact, electrons are being transferred from the zinc atoms to the hydrogen atoms (which ultimately make a molecule of diatomic hydrogen), changing the charges on both elements.

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called oxidation. The gain of electrons is called reduction. Because any loss of electrons by one substance must be accompanied by a gain in electrons by something else, oxidation and reduction always occur together. As such, electron-transfer reactions are also called oxidation-reduction reactions, or simply **redox reactions**. The atom that loses electrons is **oxidized**, and the atom that gains electrons is **reduced**. Also, because we can think of the species being oxidized as causing the reduction, the species being oxidized is called the **reducing agent**, and the species being reduced is called the **oxidizing agent**.

> *Because batteries are used as sources of electricity (that is, of electrons), all batteries are based on redox reactions.*

Although the two reactions occur together, it can be helpful to write the oxidation and reduction reactions separately as half reactions. In half reactions, we include only the reactant being oxidized or reduced, the corresponding product species, any other species needed to balance the half reaction, and the electrons being transferred. Electrons that are lost are written as products; electrons that are gained are written as reactants. For example, in our earlier equation, now written without the chloride ions,

$$\mathrm{Zn(s) + 2\,H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)}$$

zinc atoms are oxidized to Zn^{2+}. The **half reaction for the oxidation** reaction, omitting phase labels, is as follows:

$$\mathrm{Zn \rightarrow Zn^{2+} + 2\,e^-}$$

This half reaction is balanced in terms of the number of zinc atoms, and it also shows the two electrons that are needed as products to account for the zinc atom losing two negative charges to become a 2+ ion. With half reactions, there is one more item to balance: the overall charge on each side of the reaction. If you check each side of this reaction, you will note that both sides have a zero net charge.

Hydrogen is reduced in the reaction. The balanced **reduction half reaction** is as follows:

$$2\,H^{+} + 2\,e^{-} \rightarrow H_2$$

There are two hydrogen atoms on each side, and the two electrons written as reactants serve to neutralize the 2+ charge on the reactant hydrogen ions. Again, the overall charge on both sides is zero.

The overall reaction is simply the combination of the two half reactions and is shown by adding them together.

$$\begin{array}{r} Zn \rightarrow Zn^{2+} + \cancel{2e^{-}} \\ \underline{2H^{+} + \cancel{2e^{-}} \rightarrow H_2} \\ Zn + 2H^{+} \rightarrow Zn^{2+} + H_2 \end{array}$$

Because we have two electrons on each side of the equation, they can be canceled. This is the key criterion for a balanced redox reaction: the electrons have to cancel exactly. If we check the charge on both sides of the equation, we see they are the same—2+. (In reality, this positive charge is balanced by the negative charges of the chloride ions, which are not included in this reaction because chlorine does not participate in the charge transfer.)

Redox reactions are often balanced by balancing each individual half reaction and then combining the two balanced half reactions. **Sometimes a half reaction must have all of its coefficients multiplied by some integer for all the electrons to cancel**. The following example demonstrates this process.

✓ Example 1.7.4.4.1: Reducing Silver Ions

Write and balance the redox reaction that has silver ions and aluminum metal as reactants and silver metal and aluminum ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Solution

We start by using symbols of the elements and ions to represent the reaction:

$$Ag^{+} + Al \rightarrow Ag + Al^{3\,+}$$

The equation looks balanced as it is written. However, when we compare the overall charges on each side of the equation, we find a charge of +1 on the left but a charge of +3 on the right. This equation is not properly balanced. To balance it, let us write the two half reactions. Silver ions are reduced, and it takes one electron to change Ag^{+} to Ag:

Reduction half-reaction:

$$Ag^{+} + e^{-} \rightarrow Ag$$

Aluminum is oxidized, losing three electrons to change from Al to Al^{3+}:

Oxidation half-reaction:

$$Al \rightarrow Al^{3\,+} + 3\,e^{-}$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{array}{r} 3\left(Ag^{+} + e^{-} \rightarrow Ag\right) \\ \underline{Al \rightarrow Al^{3+} + 3e^{-}} \\ 3Ag^{+} + Al \rightarrow 3Ag + Al^{3+} \end{array}$$

Now the equation is balanced, not only in terms of elements but also in terms of charge.

- The **substance oxidized** is the reactant that had undergone oxidation: Al
- The **substance reduced** is the reactant that had undergone reduction: Ag^{+}
- The **reducing agent** is the same as the substance oxidized: Al
- The **oxidizing agent** is the same as the substance reduced: Ag^{+}

Exercise 1.7.4.4.1

Write and balance the redox reaction that has calcium ions and potassium metal as reactants and calcium metal and potassium ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Answer

Reduction: $Ca^{2+} + 2e^- \rightarrow Ca$

Oxidation: $2\ (K \rightarrow K^+ + e^-)$

Combined: $Ca^{2+} + 2K \rightarrow Ca + 2K^+$

- The **substance oxidized** is the reactant that had undergone oxidation: K
- The **substance reduced** is the reactant that had undergone reduction: Ca^{2+}
- The **reducing agent** is the same as the substance oxidized: K
- The **oxidizing agent** is the same as the substance reduced: Ca^{2+}

Potassium has been used as a reducing agent to obtain various metals in their elemental form.

To Your Health: Redox Reactions and Pacemaker Batteries

All batteries use redox reactions to supply electricity because electricity is basically a stream of electrons being transferred from one substance to another. Pacemakers—surgically implanted devices for regulating a person's heartbeat—are powered by tiny batteries, so the proper operation of a pacemaker depends on a redox reaction.

Pacemakers used to be powered by NiCad batteries, in which nickel and cadmium (hence the name of the battery) react with water according to this redox reaction:

$$\mathrm{Cd(s) + 2\,NiOOH(s) + 2\,H_2O(\ell) \rightarrow Cd(OH)_2(s) + 2\,Ni(OH)2\,(s)}$$

The cadmium is oxidized, while the nickel atoms in NiOOH are reduced. Except for the water, all the substances in this reaction are solids, allowing NiCad batteries to be recharged hundreds of times before they stop operating. Unfortunately, NiCad batteries are fairly heavy batteries to be carrying around in a pacemaker. Today, the lighter lithium/iodine battery is used instead. The iodine is dissolved in a solid polymer support, and the overall redox reaction is as follows:

$$\mathrm{2\,Li(s) + I_2(s) \rightarrow 2\,LiI(s)}$$

Lithium is oxidized, and iodine is reduced. Although the lithium/iodine battery cannot be recharged, one of its advantages is that it lasts up to 10 years. Thus, a person with a pacemaker does not have to worry about periodic recharging; about once per decade a person requires minor surgery to replace the pacemaker/battery unit. Lithium/iodine batteries are also used to power calculators and watches.

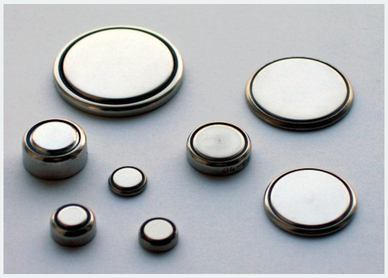


Figure 1.7.4.4.1: A small button battery like this is used to power a watch, pacemaker, or calculator. (CC BY-SA; Gerhard H Wrodnigg via Wikipedia)

Oxidation and reduction can also be defined in terms of changes in composition. The original meaning of **oxidation** was "**adding oxygen**," so when oxygen is added to a molecule, the molecule is being oxidized. The reverse is true for **reduction**: if a molecule **loses oxygen** atoms, the molecule is being reduced. For example, the acetaldehyde (CH_3CHO) molecule takes on an oxygen atom to become acetic acid (CH_3COOH).

$$\mathrm{2\,CH_3CHO + O_2 \rightarrow 2\,CH_3COOH}$$

Thus, acetaldehyde is being oxidized.

Similarly, **reduction and oxidation can be defined in terms of the gain or loss of hydrogen atoms**. If a molecule adds hydrogen atoms, it is being reduced. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of acetaldehyde into ethanol (CH_3CH_2OH), hydrogen atoms are added to acetaldehyde, so the acetaldehyde is being reduced:

$$CH_3CHO + H_2 \rightarrow CH_3CH_2OH$$

Table shows how each process affects the change in oxygen and change in hydrogen.

Process	Change in oxygen (some reactions)	Change in hydrogen (some reactions)
Oxidation	gain	lose
Reduction	lose	gain

✓ Example 1.7.4.4.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. N_2 → NH_3
b. $CH_3CH_2OHCH_3$ → CH_3COCH_3
c. HCHO → HCOOH

Solution

a. Hydrogen is being added to the original reactant molecule, so reduction is occurring.
b. Hydrogen is being removed from the original reactant molecule, so oxidation is occurring.
c. Oxygen is being added to the original reactant molecule, so oxidation is occurring.

? Exercise 1.7.4.4.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. CH_4 → CO_2 + H_2O
b. NO_2 → N_2
c. CH_2=CH_2 → CH_3CH_3

Answer a:

Oxygen is being added. Oxidation is occurring.

Answer b:

Oxygen is being removed. Reduction is occurring.

Answer a:

Hydrogen is being added. Reduction is occurring.

Key Takeaway

Chemical reactions in which electrons are transferred are called oxidation-reduction, or redox, reactions. Oxidation is the loss of electrons. Reduction is the gain of electrons. Oxidation and reduction always occur together, even though they can be written as separate chemical equations.

1.7.4.5: Practice Classifying Reactions

Classify Reactions

Exercise 1.7.4.5.1

Classify each reaction below as combination, decomposition, single displacement, double displacement, or combustion.

$CaCO_{3\ (s)} \rightarrow CO_{2\ (g)} + CaO_{\ (s)}$

Answer

decomposition

$2\ N_{2\ (g)} + 5\ O_{2\ (g)} \rightarrow 2\ N_2O_{5\ (g)}$

Answer

combination

$C_5H_{10}O_{\ (l)} + 7\ O_{2\ (g)} \rightarrow 5\ CO_{2\ (g)} + 5\ H_2O_{\ (g)}$

Answer

combustion

$ZnCl_{2\ (aq)} + 2\ AgNO_{3\ (aq)} \rightarrow Zn(NO_3)_{2\ (aq)} + 2\ AgCl_{\ (s)}$

Answer

double displacement

$Cl_{2\ (g)} + 2\ KBr_{\ (aq)} \rightarrow Br_{2\ (l)} + 2\ KCl_{\ (aq)}$

Answer

single displacement

$2\ H_{2\ (g)} + O_{2\ (g)} \rightarrow 2\ H_2O_{\ (g)}$

Answer

combination

$Mg_{\ (s)} + HCl_{\ (aq)} \rightarrow MgCl_{2\ (aq)} + H_{2\ (g)}$

Answer

single displacement

Total Ionic and Net Ionic Equations

Exercise 1.7.4.5.1

Give the total ionic equation ("complete ionic equation") and the net ionic equation for each reaction below.

$ZnCl_{2\ (aq)} + 2\ AgNO_{3\ (aq)} \rightarrow Zn(NO_3)_{2\ (aq)} + 2\ AgCl_{\ (s)}$

Total Ionic

$Zn^{2+}{}_{(aq)} + 2\ Cl^{1-}{}_{(aq)} + 2\ Ag^{1+}{}_{(aq)} + 2\ NO_3{}^{1-}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + 2\ NO_3{}^{1-}{}_{(aq)} + 2\ AgCl_{\ (s)}$

Net Ionic

$2\ Cl^{1-}{}_{(aq)} + 2\ Ag^{1+}{}_{(aq)} \rightarrow 2\ AgCl_{\ (s)}$

$H_3PO_4\ _{(aq)} + 3\ NaOH\ _{(aq)} \rightarrow Na_3PO_4\ _{(aq)} + 3\ H_2O\ _{(l)}$

Total Ionic

$3\ H^{1+}{}_{(aq)} + PO_4{}^{3-}{}_{(aq)} + 3\ Na^{1+}{}_{(aq)} + 3\ OH^{1-}{}_{(aq)} \rightarrow 3\ Na^{1+}{}_{(aq)} + PO_4{}^{3-}{}_{(aq)} + 3\ H_2O\ _{(l)}$

Net Ionic

$3\ H^{1+}{}_{(aq)} + 3\ OH^{1-}{}_{(aq)} \rightarrow 3\ H_2O\ _{(l)}$

Precipitation Reactions

Exercise 1.7.4.5.1

Which of the following compounds would not be soluble in water (and would form a solid precipitate)?

$LiCl$ $Pb(NO_3)_2$ $BaBr_2$ $CaSO_4$ $AgBr$ $Zn(C_2H_3O_2)_2$ $MgSO_4$ $CaCl_2$ K_2CO_3 $NiPO_4$

Answer

$CaSO_4$ $AgBr$ $NiPO_4$

For each of the reactions below, predict the products formed. Use the correct chemical formulas. Then balance the reaction and tell which product would be soluble ("aq" aqueous solution) and which would be insoluble ("s" solid precipitate).

$CuCl_2\ _{(aq)} + KOH\ _{(aq)} \rightarrow$

Answer

$CuCl_2\ _{(aq)} + 2\ KOH\ _{(aq)} \rightarrow 2\ KCl\ _{(aq)} + Cu(OH)_2\ _{(s)}$

$CaCl_2\ _{(aq)} + Pb(NO_3)_2\ _{(aq)} \rightarrow$

Answer

$CaCl_2\ _{(aq)} + Pb(NO_3)_2\ _{(aq)} \rightarrow Ca(NO_3)_2\ _{(aq)} + PbCl_2\ _{(s)}$

$Ba(ClO_3)_2\ _{(aq)} + Li_2SO_4\ _{(aq)} \rightarrow$

Answer

$Ba(ClO_3)_2\ _{(aq)} + Li_2SO_4\ _{(aq)} \rightarrow 2\ LiClO_3\ _{(aq)} + BaSO_4\ _{(s)}$

Oxidation and Reduction

Exercise 1.7.4.5.1

For each reaction below, tell which element is being oxidized and which is being reduced.

$Mg\ _{(s)} + HCl\ _{(aq)} \rightarrow MgCl_2\ _{(aq)} + H_2\ _{(g)}$

Answer

Mg is being oxidized (0 to +2 charge), Cl has no change (–1 to –1 charge), H is being reduced (+1 to 0 charge)

$2\ Al\ _{(s)} + 3\ Br_2\ _{(l)} \rightarrow 2\ AlBr_3\ _{(s)}$

Answer

Al is being oxidized (0 to +3 charge), Br is being reduced (0 to –1 charge)

$Cl_2\ _{(g)} + 2\ KBr\ _{(aq)} \rightarrow Br_2\ _{(l)} + 2\ KCl\ _{(aq)}$

Answer

> Cl is being reduced (0 to –1 charge), K is not changing (+1 to +1 charge), Br is being oxidized (–1 to 0 charge).

1.7.5: The Mole in Chemical Reactions

Learning Objectives

- Balance a chemical equation in terms of moles.
- Use the balanced equation to construct conversion factors in terms of moles.
- Calculate moles of one substance from moles of another substance using a balanced chemical equation.

Consider this balanced chemical equation:

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

We interpret this as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:

$$100\,H_2 + 50\,O_2 \rightarrow 100\,H_2O$$

This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:

$$5,000\,H_2 + 2,500\,O_2 \rightarrow 5,000\,H_2O$$

Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:

$$12.044 \times 10^{23} H_2 + 6.022 \times 10^{23} O_2 \rightarrow 12.044 \times 10^{23} H_2O$$

These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as "two moles of hydrogen react with one mole of oxygen to make two moles of water." All balanced chemical reactions are balanced in terms of moles.

Example 1.7.5.1

Interpret this balanced chemical equation in terms of moles.

$$P_4 + 5\,O_2 \rightarrow P_4O_{10}$$

Solution

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Exercise 1.7.5.1

Interpret this balanced chemical equation in terms of moles.

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3$$

Answer

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

In Section 4.1, we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalents—equivalents between the reactants and the products. However, now we understand that *these equivalents are expressed in terms of moles*. Consider the chemical equation

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

This chemical reaction gives us the following equivalents:

$$2 \text{ mol } H_2 \Leftrightarrow 1 \text{ mol } O_2 \Leftrightarrow 2 \text{ mol } H_2O$$

Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H_2 and 1 mol of O_2 and use it to convert from moles of one substance to moles of another:

$$17.6\ mol\,\cancel{H_2} \times \frac{1\,mol\,O_2}{2\ mol\,\cancel{H_2}} = 8.80\,mol\,O_2$$

Note how the mol H_2 unit cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation,** when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won't always be so simple!

✓ Example 1.7.5.2

For the balanced chemical equation

$$2\,C_4H_{10}(g) + 13\,O_2 \rightarrow 8\,CO_2(g) + 10\,H_2O(\ell)$$

if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is

$$13 \text{ mol } O_2 \Leftrightarrow 8 \text{ mol } CO_2$$

We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154\ mol\,\cancel{O_2} \times \frac{8\,mol\,CO_2}{13\ mol\,\cancel{O_2}} = 94.8\,mol\,CO_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they do not contribute to the number of significant figures in the final answer.

? Exercise 1.7.5.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer

118 mol

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

1.7.5.1: Mole-Mass and Mass-Mass Calculations

Learning Objectives

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Suppose we know we have 123.2 g of Cl_2. How can we determine how many moles of $AlCl_3$ we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} = 1.738\ mol\,Cl_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\,mol\ AlCl_3}{3\,mol\ Cl_2}$$

Using this conversion factor with the molar quantity we calculated above, we get

$$1.738\ \cancel{mol\,Cl_2} \times \frac{2\,mol\ AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\ AlCl_3$$

So, we will get 1.159 mol of $AlCl_3$ if we react 123.2 g of Cl_2.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2\ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} \times \frac{2\,mol\ AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\ AlCl_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 1.7.5.1.1

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of $AlCl_3$ to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of $AlCl_3$ is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$249\ \ g\,\cancel{AlCl_3} \times \frac{1\,mol\,AlCl_3}{133.33\ \ g\,\cancel{AlCl_3}} = 1.87\,mol\,AlCl_3$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of $AlCl_3$ and the number of moles of HCl:

$$frac 6\,mol\,HCl 2\ \ mol\,\cancel{AlCl_3}$$

Applying this conversion factor to the quantity of $AlCl_3$, we get

$$1.87\ \ mol\,\cancel{AlCl_3} \times \frac{6\,mol\,HCl}{2\ \ mol\,\cancel{AlCl_3}} = 5.61\,mol\,HCl$$

Alternatively, we could have done this in one line:

$$249\ \ g\,\cancel{AlCl_3} \times \frac{1\ \ mol\,\cancel{AlCl_3}}{133.33\ \ g\,\cancel{AlCl_3}} \times \frac{6\,mol\,HCl}{2\ \ mol\,\cancel{AlCl_3}} = 5.60\,mol\,HCl$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

? Exercise 1.7.5.1.1

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer

0.442 mol

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

✓ Example 1.7.5.1.2

How many grams of NH_3 will be produced when 33.9 mol of H_2 are reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have

$$33.9\ mol\,\cancel{H_2} \times \frac{2\,mol\,NH_3}{3\ mol\,\cancel{H_2}} = 22.6\,mol\,NH_3$$

Now, using the molar mass of NH_3, which is 17.03 g/mol, we get

$$22.6\ mol\,\cancel{NH_3} \times \frac{17.03\,g\,NH_3}{1\ mol\,\cancel{NH_3}} = 385\,g\,NH_3$$

Exercise 1.7.5.1.2

How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

$$\mathrm{N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g)}$$

Answer

30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO_3 that can be produced by the reaction of 45.3 g of SO_2 and O_2:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

First, we convert the given amount, 45.3 g of SO_2, to moles of SO_2 using its molar mass (64.06 g/mol):

$$45.3\ g\,\cancel{SO_2} \times \frac{1\,mol\,SO_2}{64.06\ g\,\cancel{SO_2}} = 0.707\,mol\,SO_2$$

Second, we use the balanced chemical reaction to convert from moles of SO_2 to moles of SO_3:

$$0.707\ mol\,\cancel{SO_2} \times \frac{2\,mol\,SO_3}{2\ mol\,\cancel{SO_2}} = 0.707\,mol\,SO_3$$

Finally, we use the molar mass of SO_3 (80.06 g/mol) to convert to the mass of SO_3:

$$0.707\ mol\,\cancel{SO_3} \times \frac{80.06\,g\,SO_3}{1\ mol\,\cancel{SO_3}} = 56.6\,g\,SO_3$$

We can also perform all three steps sequentially, writing them on one line as

$$45.3\ g\,\cancel{SO_2} \times \frac{1\,mol\,SO_2}{64.06\ g\,\cancel{SO_2}} \times \frac{2\,mol\,SO_3}{2\ mol\,\cancel{SO_2}} \times \frac{80.06\,g\,SO_3}{1\ mol\,\cancel{SO_3}} = 56.6\,g\,SO_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO_3, which is what we are looking for, as our final answer.

Example 1.7.5.1.3

What mass of Mg will be produced when 86.4 g of K are reacted?

$$\mathrm{MgCl_2(s) + 2\,K(s) \rightarrow Mg(s) + 2\,KCl(s)}$$

Solution

We will simply follow the steps

mass K → mol K → mol Mg → mass Mg

In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4\ g\,\cancel{K} \times \frac{1\,mol\,K}{39.09\ g\,\cancel{K}} \times \frac{1\ mol\,\cancel{Mg}}{2\ mol\,\cancel{K}} \times \frac{24.31\,g\,Mg}{1\ mol\,\cancel{Mg}} = 26.87\,g\,Mg$$

? Exercise 1.7.5.1.3

What mass of H_2 will be produced when 122 g of Zn are reacted?

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Answer

3.77 g

Summary

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

1.7.5.2: Practice Stoichiometry part 1

Mole to Mole Conversions

Exercise 1.7.5.2.1

$Fe_2O_{3\,(s)}$ + 3 $CO_{(g)}$ → 2 $Fe_{(s)}$ + 3 $CO_{2\,(g)}$

For the reaction above, how many moles of Fe would be produced from 0.877 moles Fe_2O_3?.

Answer

1.75 mol Fe

How many moles of CO_2 would be produced from 0.438 moles Fe_2O_3?.

Answer

1.31 mol CO_2

How many moles of Fe would be produced from 0.545 moles CO?.

Answer

0.363 mol Fe

Mass to Mass Conversions

Exercise 1.7.5.2.1

$Mg_3N_{2\,(s)}$ + 6 $H_2O_{(l)}$ → 3 $Mg(OH)_{2\,(s)}$ + 2 $NH_{3\,(g)}$

How many grams of H_2O would react with 34.88 grams Mg_3N_2?

Answer

$$34.88 g Mg_3N_2 \times \frac{1 mol Mg_3N_2}{100.95 g Mg_3N_2} \times \frac{6 mol H_2O}{1 mol Mg_3N_2} \times \frac{18.02 g H_2O}{1 mol H_2O}$$

= 37.36 g H_2O

How many grams of NH_3 would be produced from 76.77 grams Mg_3N_2?.

Answer

$$76.77 g Mg_3N_2 \times \frac{1 mol Mg_3N_2}{100.95 g Mg_3N_2} \times \frac{2 mol NH_3}{1 mol Mg_3N_2} \times \frac{17.04 g NH_3}{1 mol NH_3}$$

= 25.92 g NH_3

How many grams of $Mg(OH)_2$ are produced from 46.32 g H_2O?

Answer

$$46.32 g H_2O \times \frac{1 mol H_2O}{18.02 g H_2O} \times \frac{3 mol Mg(OH)_2}{6 mol H_2O} \times \frac{58.33 g Mg(OH)_2}{1 mol Mg(OH)_2}$$

= 74.97 g $Mg(OH)_2$

Mixed Stoichiometry

Exercise 1.7.5.2.1

$C_3H_{8\ (g)} + 5\ O_{2\ (g)} \rightarrow 3\ CO_{2\ (g)} + 4\ H_2O_{\ (g)}$

How many grams of O_2 are needed to produce 4.22 moles of CO_2?

Answer

225 grams of O_2

How many moles of C_3H_8 produce 87.8 grams of H_2O?

Answer

1.22 moles C_3H_8

How many grams of CO_2 are produced by a reaction that also produces 2.87 moles H_2O?

Answer

94.7 g CO_2

1.7.6: Limiting Reactant and Theoretical Yield

Learning Objectives

- Identify the limiting reactant (limiting reagent) in a given chemical reaction.
- Calculate how much product will be produced from the limiting reactant.
- Calculate how much reactant(s) remains when the reaction is complete.

In all examples discussed thus far, the reactants were assumed to be present in stoichiometric quantities. Consequently, none of the reactants were left over at the end of the reaction. This is often desirable—as in the case of a space shuttle—where excess oxygen or hydrogen is not only extra freight to be hauled into orbit, but also an explosion hazard. More often, however, reactants are present in mole ratios that are not the same as the ratio of the coefficients in the balanced chemical equation. As a result, one or more of them will not be used up completely, but will be left over when the reaction is completed. In this situation, the amount of product that can be obtained is limited by the amount of only one of the reactants. The reactant that restricts the amount of product obtained is called the limiting reactant. The reactant that remains after a reaction has gone to completion is in excess.

Consider a nonchemical example. Assume you have invited some friends for dinner and want to bake brownies for dessert. You find two boxes of brownie mix in your pantry and see that each package requires two eggs. The balanced equation for brownie preparation is:

$$1 \text{ box mix} + 2 \text{ eggs} \rightarrow 1 \text{ batch brownies} \tag{1.7.6.1}$$

If you have a dozen eggs, which ingredient will determine the number of batches of brownies that you can prepare? Because each box of brownie mix requires two eggs and you have two boxes, you need four eggs. Twelve eggs is eight more eggs than you need. Although the ratio of eggs to boxes in is 2:1, the ratio in your possession is 6:1. Hence the eggs are the ingredient (reactant) present in excess, and the brownie mix is the limiting reactant. Even if you had a refrigerator full of eggs, you could make only two batches of brownies.

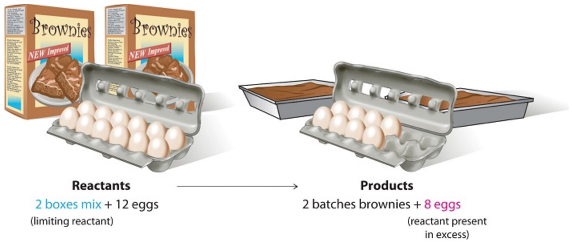


Figure 1.7.6.1: The Concept of a Limiting Reactant in the Preparation of Brownies. For a chemist, the **balanced** chemical equation is the recipe that must be followed. 2 boxes of brownie mix and 12 eggs results in 2 batches of brownies and 8 eggs; in this case the 8 eggs are reactant present in excess

PhET Simulation: Reactants, Products and Leftovers

View this interactive simulation illustrating the concepts of limiting and excess reactants.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$H_2 + Cl_2(g) \rightarrow 2\,HCl(g)$$

The balanced equation shows that hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen non-reacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example, in the previous paragraph, complete reaction of the hydrogen would yield:

$$\text{mol HCl produced} = 3 \text{ mol } H_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2} = 6 \text{ mol HCl}$$

Complete reaction of the provided chlorine would produce:

$$\text{mol HCl produced} = 2 \text{ mol } Cl_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Cl_2} = 4 \text{ mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure 1.7.6.2).

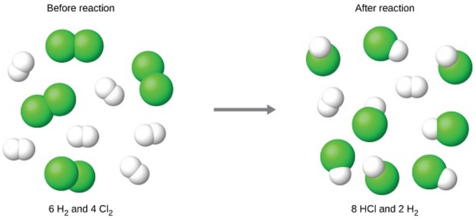


Figure 1.7.6.2: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant. The figure shows a space-filling molecular models reacting. There is a reaction arrow pointing to the right in the middle. To the left of the reaction arrow there are three molecules each consisting of two green spheres bonded together. There are also five molecules each consisting of two smaller, white spheres bonded together. Above these molecules is the label, "Before reaction," and below these molecules is the label, "6 H subscript 2 and 4 C l subscript 2." To the right of the reaction arrow, there are eight molecules each consisting of one green sphere bonded to a smaller white sphere. There are also two molecules each consisting of two white spheres bonded together. Above these molecules is the label, "After reaction," and below these molecules is the label, "8 H C l and 2 H subscript 2."

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reactant; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reactant and which is in excess.

How to Identify the Limiting Reactant (Limiting Reagent)

There are two ways to determine the limiting reactant. One method is to find and compare the mole ratio of the reactants used in the reaction (Approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the **smallest** amount of product is the limiting reactant (Approach 2). This section will focus more on the second method.

Approach 1 (The "Reactant Mole Ratio Method"): Find the limiting reactant by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting (excess) reactant.

Approach 2 (The "The Product Method"): Find the limiting reactant by calculating and comparing the amount of product that each reactant will produce.

1. Balance the chemical equation for the chemical reaction.
2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reactant.
5. The reactant that produces a larger amount of product is the excess reactant.
6. To find the amount of remaining excess reactant, subtract the mass of excess reactant consumed from the total mass of excess reactant given.

The key to recognizing which reactant is the limiting reactant is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reactant. What we need to do is determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reactant. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

Example 1.7.6.1: Identifying the Limiting Reactant

As an example, consider the balanced equation

$$4\,C_2H_3Br_3 + 11\,O_2 \rightarrow 8\,CO_2 + 6\,H_2O + 6\,Br_2$$

What is the limiting reactant if 76.4 grams of $C_2H_3Br_3$ reacted with 49.1 grams of O_2?

Solution

Using Approach 1:

Step 1: *Balance the chemical equation.*

The equation is already balanced with the relationship

$$4 \text{ mol } C_2H_3Br_3 \text{ to } 11 \text{ mol } O_2 \text{ to } 6 \text{ mol } H_2O \text{ to } 6 \text{ mol Br}$$

Step 2: Convert all given information into moles.

$$76.4 \;\; \cancel{\text{g } C_2H_3Br_3} \times \frac{1 \text{ mol } C_2H_3Br_3}{266.72 \;\; \cancel{\text{g } C_2H_3B_3}} = 0.286 \text{ mol } C_2H_3Br_3$$

$$49.1 \;\; \cancel{\text{g } O_2} \times \frac{1 \text{ mol } O_2}{32.00 \;\; \cancel{\text{g } O_2}} = 1.53 \text{ mol } O_2$$

Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.

Assuming that all of the oxygen is used up,

$$1.53\ \ \text{mol}\cancel{O_2} \times \frac{4\ \text{mol}C_2H_3Br_3}{11\ \ \text{mol}\cancel{O_2}} = 0.556 \text{ mol } C_2H_3Br_3 \text{ are required.}$$

Because 0.556 moles of $C_2H_3Br_3$ required > 0.286 moles of $C_2H_3Br_3$ available, $C_2H_3Br_3$ is the limiting reactant.

Using Approach 2:

Step 1: *Balance the chemical equation.*

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol O_2 to 6 mol H_2O to 6 mol Br_2

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$76.4\ \ \text{g}\ \cancel{C_2H_3Br_3} \times \frac{1\ \ \text{mol}\ \cancel{C_2H_3Br_3}}{266.72\ \ \text{g}\ \cancel{C_2H_3Br_3}} \times \frac{8\ \ \text{mol}\ \cancel{CO_2}}{4\ \ \text{mol}\ \cancel{C_2H_3Br_3}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \ \text{mol}\ \cancel{CO_2}} = 25.2\ \text{g}\ CO_2$$

$$49.1\ \ \text{g}\ \cancel{O_2} \times \frac{1\ \ \text{mol}\ \cancel{O_2}}{32.00\ \ \text{g}\ \cancel{O_2}} \times \frac{8\ \ \text{mol}\ \cancel{CO_2}}{11\ \ \text{mol}\ \cancel{O_2}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \ \text{mol}\ \cancel{CO_2}} = 49.1\ \text{g}\ CO_2$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Therefore, by either method, $C_2H_3Br_3$ is the limiting reactant.

Example 1.7.6.2: Identifying the Limiting Reactant and the Mass of Excess Reactant

For example, in the reaction of magnesium metal and oxygen, calculate the mass of magnesium oxide that can be produced if 2.40 g Mg reacts with 10.0 g O_2. Also determine the amount of excess reactant. MgO is the only product in the reaction.

Solution

Following Approach 1:

Step 1: Balance the chemical equation.

2 Mg (s) + O_2 (g) → **2** MgO (s)

The balanced equation provides the relationship of 2 mol Mg to 1 mol O_2 to 2 mol MgO

Step 2 and Step 3: Convert mass to moles and stoichiometry.

$$2.40\ \ \text{g}\ \cancel{\text{Mg}} \times \frac{1\ \ \text{mol}\ \cancel{\text{Mg}}}{24.31\ \ \text{g}\ \cancel{\text{Mg}}} \times \frac{2\ \ \text{mol}\ \cancel{\text{MgO}}}{2\ \ \text{mol}\ \cancel{\text{Mg}}} \times \frac{40.31\ \text{g}\ \text{MgO}}{1\ \ \text{mol}\ \cancel{\text{MgO}}} = 3.98\ \text{g}\ \text{MgO}$$

$$10.0\ \ \text{g}\ \cancel{O_2} \times \frac{1\ \ \text{mol}\ \cancel{O_2}}{32.00\ \ \text{g}\ \cancel{O_2}} \times \frac{2\ \ \text{mol}\ \cancel{\text{MgO}}}{1\ \ \text{mol}\ \cancel{O_2}} \times \frac{40.31\ \text{g}\ \text{MgO}}{1\ \ \text{mol}\ \cancel{\text{MgO}}} = 25.2\ \text{g}\ \text{MgO}$$

Step 4: The reactant that produces a smaller amount of product is the limiting reactant.

Mg produces less MgO than does O_2 (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reactant in this reaction.

Step 5: The reactant that produces a larger amount of product is the excess reactant.

O_2 produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore O_2 is the excess reactant in this reaction.

Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reactant consumed from the total mass of excess reactant given.

Mass of excess reactant calculated using the limiting reactant:

$$2.40\ \ \text{g}\ \cancel{\text{Mg}} \times \frac{1\ \ \text{mol}\ \cancel{\text{Mg}}}{24.31\ \ \text{g}\ \cancel{\text{Mg}}} \times \frac{1\ \ \text{mol}\ \cancel{O_2}}{2\ \ \text{mol}\ \cancel{\text{Mg}}} \times \frac{32.00\ \text{g}\ O_2}{1\ \ \text{mol}\ \cancel{O_2}} = 1.58\ \text{g}\ O_2$$

OR

Mass of excess reactant calculated using the mass of the product:

$$3.98\ \ \text{g}\ \cancel{\text{MgO}} \times \frac{1\ \ \text{mol}\ \cancel{\text{MgO}}}{40.31\ \ \text{g}\ \cancel{\text{MgO}}} \times \frac{1\ \ \text{mol}\ \cancel{O_2}}{2\ \ \text{mol}\ \cancel{\text{MgO}}} \times \frac{32.0\ \text{g}\ O_2}{1\ \ \text{mol}\ \cancel{O_2}} = 1.58\ \text{g}\ O_2$$

Mass of total excess reactant given – mass of excess reactant consumed in the reaction:

10.0g O_2 - (available) 1.58g O_2 (used) = 8.42g O_2 (excess)

Therefore, O_2 is in excess.

Example 1.7.6.3: Limiting Reactant

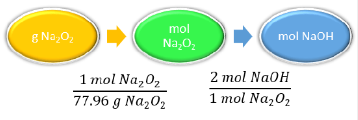
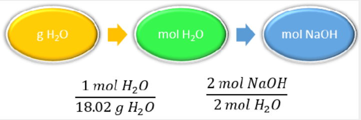
What is the limiting reactant if 78.0 grams of Na_2O_2 were reacted with 29.4 grams of H_2O? The unbalanced chemical equation is

$$Na_2O_2(s) + H_2O(l) \rightarrow NaOH(aq) + H_2O_2(l)$$

Solution

Solutions to Example 8.4.3

Steps for Problem Solving- The Product Method	Example 1.7.6.1
Identify the "given" information and what the problem is asking you to "find."	Given: 78.0 grams of Na_2O_2 29.4 g H_2O Find: limiting reactant
List other known quantities.	1 mol Na_2O_2= 77.96 g/mol 1 mol H_2O = 18.02 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.

Steps for Problem Solving- The Product Method	Example 1.7.6.1
Balance the equation.	Na_2O_2 (s) + **2**H_2O (l) → **2**NaOH (aq) + H_2O_2 (l) The balanced equation provides the relationship of 1 mol Na_2O_2 to 2 mol H_2O 2mol NaOH to 1 mol H_2O_2
Prepare a concept map and use the proper conversion factor.	  Because the question only asks for the limiting reactant, we can perform two **mass-mole** calculations and determine which amount is less.
Cancel units and calculate.	$78.0\text{ g Na}_2\text{O}_2 \times \frac{1\text{ mol Na}_2\text{O}_2}{77.96\text{ g Na}_2\text{O}_2} \times \frac{2\text{ mol NaOH}}{1\text{ mol Na}_2\text{O}_2} \times \frac{40\text{ g NaOH}}{1\text{ mol NaOH}} = 2.00\text{ mol NaOH}$ $29.4\text{ g H}_2\text{O} \times \frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} \times \frac{2\text{ mol NaOH}}{2\text{ mol Na}_2\text{O}_2} \times \frac{40\text{ g NaOH}}{1\text{ mol NaOH}} = 1.63\text{ mol NaOH}$ Therefore, H_2O is the *limiting reactant*.

Think about your result.

Example 1.7.6.4: Limiting Reactant and Mass of Excess Reactant

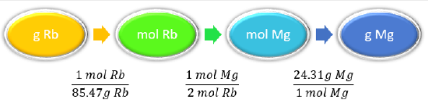
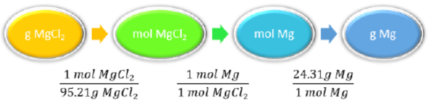
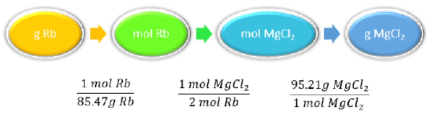
A 5.00 g quantity of Rb is combined with 3.44 g of MgCl_2 according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of Mg is formed, and what mass of remaining reactant is left over?

Solution

Solutions to Example 8.4.10

Steps for Problem Solving- The Product Method	Example 1.7.6.2
Steps for Problem Solving	A 5.00 g quantity of Rb is combined with 3.44 g of MgCl2 according to this chemical reaction: $2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$ What mass of Mg is formed, and what mass of remaining reactant is left over?
Identify the "given" information and what the problem is asking you to "find."	Given: 5.00g Rb, 2.44g MgCl2 Find: mass of Mg formed, mass of remaining reactant
List other known quantities.	• molar mass: Rb = 85.47 g/mol • molar mass: $MgCl_2$ = 95.21 g/mol • molar mass: Mg = 24.31 g/mol
Prepare concept maps and use the proper conversion factor.	Find mass Mg formed based on mass of Rb  Find mass of Mg formed based on mass of $MgCl_2$  Conversion factors: 1 mole MgCl2 to 95.21 grams MgCl2, 1 mole Mg to 1 mole MgCl2, 24.31 grams Mg to 1 mole Mg Use limiting reactant to determine amount of excess reactant consumed  Conversion factors: 1 mole Rb to 85.47 grams Rb, 1 mole MgCl2 to 2 moles Rb, 95.21 grams MgCl2 to 1 mole MgCl2

Steps for Problem Solving- The Product Method	Example 1.7.6.2
Cancel units and calculate.	Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less. $5.00\ g\,\cancel{Rb} \times \frac{1\ mol\,\cancel{Rb}}{85.47\ g\,\cancel{Rb}} \times \frac{1\ mol\,\cancel{Mg}}{2\ mol\,\cancel{Rb}} \times \frac{24.31\ g\,Mg}{1\ mol\,\cancel{Mg}} = 0.711\ g\,Mg$ $3.44\ g\,\cancel{MgCl_2} \times \frac{1\ mol\,\cancel{MgCl_2}}{95.21\ g\,\cancel{MgCl_2}} \times \frac{1\ mol\,\cancel{Mg}}{1\ mol\,\cancel{MgCl_2}} \times \frac{24.31\ g\,Mg}{1\ mol\,\cancel{Mg}} = 0.878\ g\,Mg$ The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reactant. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of $MgCl_2$ reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount. $5.00\ g\,\cancel{Rb} \times \frac{1\ mol\,\cancel{Rb}}{85.47\ g\,\cancel{Rb}} \times \frac{1\ mol\,\cancel{MgCl_2}}{2\ mol\,\cancel{Rb}} \times \frac{95.21\ g\,MgCl_2}{1\ mol\,\cancel{MgCl_2}} = 2.78\ g\,MgCl_2\ \ reacted$ Because we started with 3.44 g of $MgCl_2$, we have 3.44 g $MgCl_2$ − 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ left
Think about your result.	It usually is not possible to determine the limiting reactant using just the initial masses, as the reagents have different molar masses and coefficients.

Exercise 1.7.6.1

Given the initial amounts listed, what is the limiting reactant, and what is the mass of the leftover reactant?

$$\underbrace{22.7\,g}_{MgO(s)} + \underbrace{17.9\,g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer

H_2S is the limiting reagent; 1.5 g of MgO are left over.

1.7.6.1: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants

Learning Objectives

- Calculate percentage or actual yields from known amounts of reactants.

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

Percent Yield

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, the maximum amount of product that can be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

Typically, percent yields are understandably less than 100% because of the reasons indicated earlier. However, percent yields greater than 100% are possible if the measured product of the reaction contains impurities that cause its mass to be greater than it actually would be if the product was pure. When a chemist synthesizes a desired chemical, he or she is always careful to purify the products of the reaction. Example $1.7.6.1.1$ illustrates the steps for determining percent yield.

Example 1.7.6.1.1: Decomposition of Potassium Chlorate

Potassium chlorate decomposes upon slight heating in the presence of a catalyst, according to the reaction below:

$$2\text{KClO}_3\,(s) \rightarrow 2\text{KCl}\,(s) + 3\text{O}_2\,(g)$$

In a certain experiment, 40.0 g KClO_3 is heated until it completely decomposes. The experiment is performed and the oxygen gas is collected and its mass is found to be 14.9 g.

a. What is the theoretical yield of oxygen gas?
b. What is the percent yield for the reaction?

Solution

a. Calculation of theoretical yield

First, we will calculate the theoretical yield based on the stoichiometry.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Mass of $\text{KClO}_3 = 40.0 \text{ g}$

Mass of O_2 collected = 14.9g

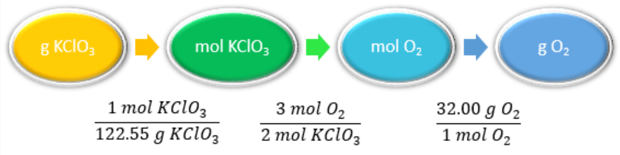
Find: Theoretical yield, g O_2

Step 2: List other known quantities and plan the problem.

1 mol $KClO_3$ = 122.55 g/mol

1 mol O_2 = 32.00 g/mol

Step 3: Apply stoichiometry to convert from the mass of a reactant to the mass of a product:



Step 4: Solve.

$$40.0 \text{ g } \cancel{\text{KClO}_3} \times \frac{1 \text{ mol } \cancel{\text{KClO}_3}}{122.55 \text{ g } \cancel{\text{KClO}_3}} \times \frac{3 \text{ mol } \cancel{\text{O}_2}}{2 \text{ mol } \cancel{\text{KClO}_3}} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol } \cancel{\text{O}_2}} = 15.7 \text{ g O}_2$$

The theoretical yield of O_2 is 15.7 g, $15.\underline{6}7$ g unrounded.

Step 5: Think about your result.

The mass of oxygen gas must be less than the 40.0 g of potassium chlorate that was decomposed.

b. Calculation of percent yield

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

Step 1: Identify the "given" information and what the problem is asking you to "find".

Given: Theoretical yield =$15.\underline{6}7$ g, use the un-rounded number for the calculation.

Actual yield = 14.9g

Find: Percent yield, % Yield

Step 2: List other known quantities and plan the problem.

No other quantities needed.

Step 3: Use the percent yield equation below.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Step 4: Solve.

$$\text{Percent Yield} = \frac{14.9 \text{ g}}{15.\underline{6}7 \text{ g}} \times 100\% = 94.9\%$$

Step 5: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

✓ Example 1.7.6.1.2: Oxidation of Zinc

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

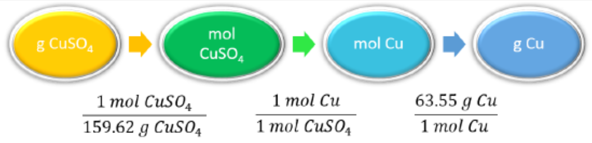
$$\text{CuSO}_4(aq) + \text{Zn}(s) \rightarrow \text{Cu}(s) + \text{ZnSO}_4(aq)$$

What is the percent yield?

Solution

Solutions to Example 8.6.2

Steps for Problem Solving-The Product Method	Example 1.7.6.1.1
Identify the "given" information and what the problem is asking you to "find."	Given: 1.274 g CuSO4 Actual yield = 0.392 g Cu Find: Percent yield
List other known quantities.	1 mol CuSO4= 159.62 g/mol 1 mol Cu = 63.55 g/mol Since the amount of product in grams is not required, only the molar mass of the reactants is needed.
Balance the equation.	The chemical equation is already balanced. The balanced equation provides the relationship of 1 mol CuSO4 to 1 mol Zn to 1 mol Cu to 1 mol ZnSO4.

Steps for Problem Solving-The Product Method	Example $1.7.6.1.1$
Prepare a concept map and use the proper conversion factor.	 The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield (g Cu) is found by performing **mass-mass** calculation based on the initial amount of $CuSO_4$.
Cancel units and calculate.	$1.274\ \text{g}\ \cancel{\text{CuSO}_4} \times \frac{1\ \text{mol}\ \cancel{\text{CuSO}_4}}{159.62\ \text{g}\ \cancel{\text{CuSO}_4}} \times \frac{1\ \text{mol}\ \cancel{\text{Cu}}}{1\ \text{mol}\ \cancel{\text{CuSO}_4}} \times \frac{63.55\ \text{g Cu}}{1\ \text{mol}\ \cancel{\text{Cu}}} = 0.507$ Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be: $\begin{aligned} \text{percent yield} &= \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100 \\ &= \left(\frac{0.392\, g\, \text{Cu}}{0.5072\, g\, \text{Cu}}\right) \times 100 \\ &= 77.3\% \end{aligned}$
Think about your result.	Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

? Exercise $1.7.6.1.1$

What is the percent yield of a reaction that produces 12.5 g of the Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$\text{CCl}_4 + 2\,\text{HF} \rightarrow \text{CF}_2\text{Cl}_2 + 2\,\text{HCl}$$

Answer

48.3%

Summary

Theoretical yield is calculated based on the stoichiometry of the chemical equation. The actual yield is experimentally determined. The percent yield is determined by calculating the ratio of actual yield to theoretical yield.

1.7.6.2: Practice Stoichiometry part 2

Exercise 1.7.6.2.1

$2\ Mg_{\ (s)} + O_{2\ (g)} \rightarrow 2\ MgO_{\ (s)}$

In the reaction above, how many grams of MgO would be produced from 6.87 g of Mg and 5.22 g O_2?

Answer

11.4 g MgO (Mg is limiting)

Exercise 1.7.6.2.1

$2\ NH_{3\ (g)} + 3\ O_{2\ (g)} \rightarrow 3\ H_2O_{\ (g)} + N_2O_{3\ (g)}$

In the reaction above, how many grams of N_2O_3 would be produced from 12.76 g of NH_3 and 17.44 g O_2?

Answer

13.81 g N_2O_3 (O_2 is limiting)

In the reaction above, how many grams of N_2O_3 would be produced from 0.765 moles NH_3 and 1.206 moles O_2?

Answer

29.1 g (NH_3 is limiting)

Exercise 1.7.6.2.1

$2\ HgO_{\ (s)} \rightarrow 2\ Hg_{\ (l)} + O_{2\ (g)}$

If you heat 375.0 grams of HgO and obtain 248.4 g Hg, what is your percent yield?

Answer

71.52 % (expect 347.3 g Hg product)

Exercise 1.7.6.2.1

$6\ Li_{\ (s)} + N_{2\ (g)} \rightarrow 2\ Li_3N_{\ (s)}$

If you heat 38.44 grams of Li (with plenty of nitrogen) and obtain 44.74 g Li_3N, what is your percent yield?

Answer

69.57 % (expect 64.31 g Li_3N product)

Exercise 1.7.6.2.1

$Fe_2S_{3\ (s)} + 6\ HCl_{\ (g)} \rightarrow 2\ FeCl_{3\ (s)} + 3\ H_2S_{\ (g)}$

If you treat 17.82 grams of Fe_2S_3 (with plenty of HCl) and obtain 21.36 g $FeCl_3$, what is your percent yield?

Answer

76.83 % (expect 27.80 g $FeCl_3$ product)

1.7.7: Enthalpy Change is a Measure of the Heat Evolved or Absorbed

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe, called the system and the surroundings. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** are everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings is the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

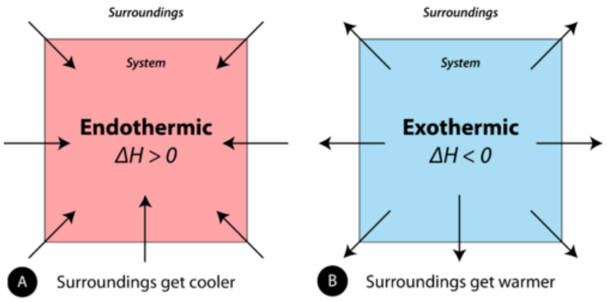


Figure 1.7.7.1: (A) Endothermic reaction. (B) Exothermic reaction.

Enthalpy

Heat changes in chemical reactions are often measured in the laboratory under conditions in which the reacting system is open to the atmosphere. In that case, the system is at a constant pressure. **Enthalpy** $(\boldsymbol{H})$ is the heat content of a system at constant pressure. Chemists routinely measure changes in enthalpy of chemical systems as reactants are converted into products. The heat that is absorbed or released by a reaction at constant pressure is the same as the enthalpy change, and is given the symbol ΔH. Unless otherwise specified, all reactions in this material are assumed to take place at constant pressure.

The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products. The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds. Energy needs to be put into the system in order to break chemical bonds, as they do not come apart spontaneously in most cases. Bond formation to produce products will involve release of energy. The change in enthalpy shows the trade-offs made in these two processes. Does it take more energy to break bonds than that needed to form bonds? If so, the reaction is endothermic and the enthalpy change is positive. If more energy is produced in bond formation than that needed for bond breaking, the reaction is exothermic and the enthalpy is negative.

Several factors influence the enthalpy of a system. Enthalpy is an extensive property, determined in part by the amount of material we work with. The state of reactants and products (solid, liquid, or gas) influences the enthalpy value for a system. The direction of the reaction affects the enthalpy value. A reaction that takes place in the opposite direction has the same numerical enthalpy value, but the opposite sign.

Thermochemical Equation

When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of 1 mol of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of the balanced equation:

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) + 890.4\text{ kJ}$$

The equation tells us that 1 mol of methane combines with 2 mol of oxygen to produce 1 mol of carbon dioxide and 2 mol of water. In the process, 890.4 kJ is released and so it is written as a product of the reaction. A **thermochemical equation** is a chemical equation that includes the enthalpy change of the reaction. The process in the above thermochemical equation can be shown visually in Figure 1.7.7.2

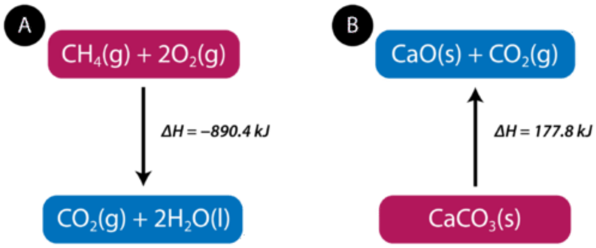


Figure 1.7.7.2: (A) As reactants are converted to products in an exothermic reaction, enthalpy is released into the surroundings. The enthalpy change of the reaction is negative. (B) As reactants are converted to products in an endothermic reaction, enthalpy is absorbed from the surroundings. The enthalpy change of the reaction is positive.

In the combustion of methane example, the enthalpy change is negative because heat is being released by the system. Therefore, the overall enthalpy of the system decreases. The **heat of reaction** is the enthalpy change for a chemical reaction. In the case above, the heat of reaction is -890.4 kJ. The thermochemical reaction can also be written in this way:

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) \qquad \Delta H = -890.4\ \mathrm{kJ}$$

Heats of reaction are typically measured in kilojoules. It is important to include the physical states of the reactants and products in a thermochemical equation as the value of the ΔH depends on those states.

Endothermic reactions absorb energy from the surroundings as the reaction occurs. When 1 mol of calcium carbonate decomposes into 1 mol of calcium oxide and 1 mol of carbon dioxide, 177.8 kJ of heat is absorbed. The process is shown visually in Figure 1.7.7.2B. The thermochemical reaction is shown below.

$$\mathrm{CaCO_3}\,(s) + 177.8\ \mathrm{kJ} \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g)$$

Because the heat is absorbed by the system, the 177.8 kJ is written as a reactant. The heat of reaction is positive for an endothermic reaction.

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \qquad \Delta H = 177.8\ \mathrm{kJ}$$

The way in which a reaction is written influences the value of the enthalpy change for the reaction. Many reactions are reversible, meaning that the product(s) of the reaction are capable of combining and reforming the reactant(s). If a reaction is written in the reverse direction, the sign of the ΔH changes. For example, we can write an equation for the reaction of calcium oxide with carbon dioxide to form calcium carbonate.

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) + 177.8\ \mathrm{kJ}$$

The reaction is exothermic and thus the sign of the enthalpy change is negative.

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) \qquad \Delta H = -177.8\ \mathrm{kJ}$$

Stoichiometric Calculations and Enthalpy Changes

Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems. Refer again to the combustion reaction of methane. Since the reaction of 1 mol of methane released 890.4 kJ, the reaction of 2 mol of methane would release $2 \times 890.4\ \mathrm{kJ} = 1781\ \mathrm{kJ}$. The reaction of 0.5 mol of methane would release $\frac{890{,}4\ \mathrm{kJ}}{2} = 445.2\ \mathrm{kJ}$. As with other stoichiometry problems, the moles of a reactant or product can be linked to mass or volume.

✓ Example 1.7.7.1

Sulfur dioxide gas reacts with oxygen to form sulfur trioxide in an exothermic reaction, according to the following thermochemical equation.

$$2\mathrm{SO_2}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{SO_3}\,(g) + 198\ \mathrm{kJ}$$

Calculate the enthalpy change that occurs when $58.0 \mathrm{~g}$ of sulfur dioxide is reacted with excess oxygen.

Solution

Step 1: List the known quantities and plan the problem.

- Mass $\mathrm{SO_2} = 58.0 \mathrm{~g}$
- Molar mass $\mathrm{SO_2} = 64.07 \mathrm{~g/mol}$
- $\Delta H = -198 \mathrm{~kJ}$ for the reaction of $2 \mathrm{~mol~SO_2}$

Unknown

- $\Delta H = ? \mathrm{~kJ}$

The calculation requires two steps. The mass of $\mathrm{SO_2}$ is converted to moles. Then the moles of $\mathrm{SO_2}$ is multiplied by the conversion factor of $\left(\frac{-198 \mathrm{~kJ}}{2 \mathrm{~mol~SO_2}}\right)$.

Step 2: Solve.

$$\Delta H = 58.0 \mathrm{~g~SO_2} \times \frac{1 \mathrm{~mol~SO_2}}{64.07 \mathrm{~g~SO_2}} \times \frac{-198 \mathrm{~kJ}}{2 \mathrm{~mol~SO_2}} = 89.6 \mathrm{~kJ}$$

Step 3: Think about your result.

The mass of sulfur dioxide is slightly less than $1 \mathrm{~mol}$. Since $198 \mathrm{~kJ}$ is released for every $2 \mathrm{~mol}$ of $\mathrm{SO_2}$ that reacts, the heat released when about $1 \mathrm{~mol}$ reacts is one half of 198. The $89.6 \mathrm{~kJ}$ is slightly less than half of 198. The sign of ΔH is negative because the reaction is exothermic.

1.7.7.1: Practice Stoichiometry Part 3

Exercise 1.7.7.1.1

2 NO $_{(g)}$ + O$_{2\ (g)}$ $\rightarrow$ 2 NO$_{2\ (g)}$ $\Delta H =$ – 114 kJ

Is the reaction above endothermic or exothermic? What is the energy change when 0.633 moles of NO react?

Answer

Exothermic. Energy change is – 36.1 kJ, meaning that 36.1 kJ of energy is released.

Exercise 1.7.7.1.1

CaCO$_{3\ (s)}$ $\rightarrow$ CO$_{2\ (g)}$ + CaO $_{(s)}$ $\Delta H =$ + 192 kJ

Is the reaction above endothermic or exothermic? What is the energy change when 2.87 grams of $CaCO_3$ react?

Answer

Endothermic. Energy change is + 5.51 kJ, meaning that 5.51 kJ of energy is absorbed.

Exercise 1.7.7.1.1

S$_{8\ (s)}$ + 12 O$_{2\ (g)}$ $\rightarrow$ 8 SO$_{3\ (g)}$ $\Delta H =$ – 3166 kJ

Is the reaction above endothermic or exothermic? What is the energy change when 26.7 grams of SO_3 are produced?

Answer

Exothermic. Energy change is – 132 kJ, meaning that 132 kJ of energy is released.

Exercise 1.7.7.1.1

2 HCl $_{(g)}$ $\rightarrow$ H$_{2\ (g)}$ + Cl$_{2\ (g)}$ $\Delta H =$ + 185 kJ

Is the reaction above endothermic or exothermic? What is the energy change when 489 grams of HCl react?

Answer

Endothermic. Energy change is + 1240 kJ, meaning that 1240 kJ of energy is absorbed.

SECTION OVERVIEW

1.8: Gases

1.8.1: Kinetic Theory of Gases

Learning Objectives

- State the major concepts behind the kinetic theory of gases.
- Relate the general properties of gases to the kinetic theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. The kinetic theory of gases is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic—that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas.

Figure 1.8.1.1shows a representation of how we mentally picture the gas phase.

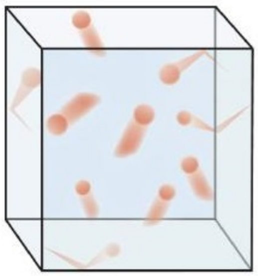


Figure 1.8.1.1: The Kinetic Theory of Gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion, with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements, and the kinetic theory of gases is well accepted by the scientific community.

- The physical behavior of gases is explained by the kinetic theory of gases.
- An ideal gas adheres exactly to the kinetic theory of gases.

1.8.2: Pressure- The Result of Constant Molecular Collisions

Learning Objectives

- Define *pressure*.
- Learn the units of pressure and how to convert between them.

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being “held in” by the earth’s gravity, rather than the gas being in a container. The pressure of the atmosphere is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

However, “average atmospheric pressure at sea level” is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mmHg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalents:

1 atm=760 mmHg=760 torr

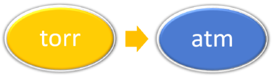
We can use these equivalents as with any equivalence—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, 1 atm = 101,325 Pa.

Example 1.8.2.1: Pressure Conversion

How many atmospheres are there in 595 torr?

Solution

Solutions to Example 11.3.1

Steps for Problem Solving	Unit Conversion
Identify the "given” information and what the problem is asking you to "find."	Given: 595 torr Find: ? atm
List other known quantities.	1 atm = 760 torr
Prepare a concept map.	 $\frac{1\ atm}{760\ torr}$
Cancel units and calculate.	$595\ \cancel{torr} \times \frac{1\ atm}{760\ \cancel{torr}} = 0.783\ atm$
Think about your result.	595 torr is less than 760 torr so the final answer should be less than 1 atm.

? Exercise 1.8.2.1

How many atmospheres are there in 1,022 torr?

Answer

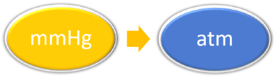
1.345 atm

✓ Example 1.8.2.2: Mars

The atmosphere on Mars is largely CO_2 at a pressure of 6.01 mmHg. What is this pressure in atmospheres?

Solution

Solutions to Example 11.3.2

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 6.01mmHg Find: ? atm
List other known quantities.	1 atm = 760 mmHg
Prepare a concept map.	 $\frac{1\ atm}{760\ mmHg}$
Cancel units and calculate.	$6.01\ \cancel{mmHg} \times \frac{1\ atm}{760\ \cancel{mmHg}} = 0.00791\ atm = 7.91 \times 10^{-3} atm$
Think about your result.	6.01 is a very small number relative to 760 mmHg, just like the value in atmospheres.

? Exercise 1.8.2.2

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer

652 torr

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

1.8.2.1: Practice Pressure Units

Exercise 1.8.2.1.1

The cabin on a passenger airplane at 30,000 feet is "pressurized" so passengers experience 586 mm Hg of air pressure. This is necessary because outside the air pressure is only 226 mm Hg, and your body can't get enough oxygen at that pressure! What are these pressures in atm? (Remember 1 atm = 760 mm Hg exactly)

Answer

Inside = 0.771 atm; outside = 0.297 atm.

Exercise 1.8.2.1.1

The pressure limit on a class D oxygen tank is 4000.0 psi. What pressure is that in kPa? Some relationships are listed below.

1.00000 atm = 101.325 kPa = 14.6959 psi

Answer

27579 kPa

1.8.3: Boyle's Law

Each day, hundreds of weather balloons are launched. Made of a synthetic rubber and carrying a box of instruments, each helium-filled balloon rises up into the sky. As a balloon gains altitude, the atmospheric pressure becomes less and the balloon expands. At some point the balloon bursts due to the expansion; the instruments drop (aided by a parachute) to be retrieved and studied for information about the weather.

Boyle's Law

Robert Boyle (1627-1691), an English chemist, is widely considered to be one of the founders of the modern experimental science of chemistry. He discovered that doubling the pressure of an enclosed sample of gas, while keeping its temperature constant, caused the volume of the gas to be reduced by half. **Boyle's law** states that the volume of a given mass of gas varies inversely with the pressure when the temperature is kept constant. An inverse relationship is described in this way. As one variable increases in value, the other variable decreases.

Physically, what is happening? The gas molecules are moving and are a certain distance apart from one another. An increase in pressure pushes the molecules closer together, reducing the volume. If the pressure is decreased, the gases are free to move about in a larger volume.

Figure 1.8.3.1: Robert Boyle. (CC BY-NC; CK-12)

Mathematically, Boyle's law can be expressed by the equation:

$$P \times V = k$$

The k is a constant for a given sample of gas and depends only on the mass of the gas and the temperature. The table below shows pressure and volume data for a set amount of gas at a constant temperature. The third column represents the value of the constant (k) for this data and is always equal to the pressure multiplied by the volume. As one of the variables changes, the other changes in such a way that the product of $P \times V$ always remains the same. In this particular case, that constant is $500 \text{ atm} \cdot \text{mL}$.

Table 1.8.3.1: Pressure-Volume Data

Pressure (atm)	**Volume** (mL)	$\boldsymbol{P \times V = k}$ $(\text{atm} \cdot \text{mL})$
0.5	1000	500
0.625	800	500
1.0	500	500
2.0	250	500
5.0	100	500
8.0	62.5	500
10.0	50	500

A graph of the data in the table further illustrates the inverse relationship nature of Boyle's Law (see figure below). Volume is plotted on the x-axis, with the corresponding pressure on the y-axis.

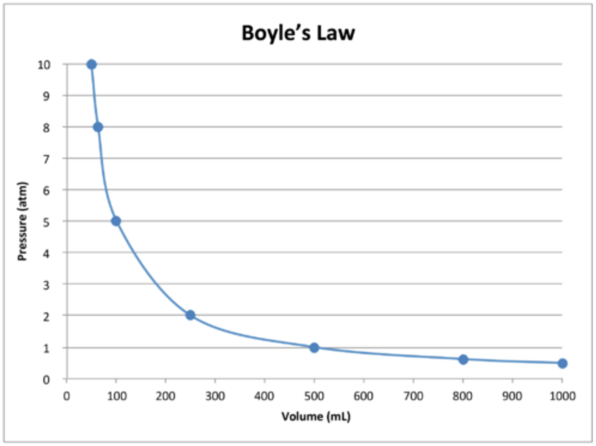


Figure 1.8.3.2: The pressure of a gas decreases as the volume increases, making Boyle's law an inverse relationship. (CC BY-NC; CK-12)

Boyle's Law can be used to compare changing conditions for a gas. We use P_1 and V_1 to stand for the initial pressure and initial volume of a gas. After a change has been made, P_2 and V_2 stand for the final pressure and volume. The mathematical relationship of Boyle's Law becomes:

$$P_1 \times V_1 = P_2 \times V_2$$

This equation can be used to calculate any one of the four quantities if the other three are known.

Example 1.8.3.1

A sample of oxygen gas has a volume of 425 mL when the pressure is equal to 387 kPa. The gas is allowed to expand into a 1.75 L container. Calculate the new pressure of the gas.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P_1 = 387 \text{ kPa}$
- $V_1 = 425 \text{ mL}$
- $V_2 = 1.75 \text{ L} = 1750 \text{ mL}$

Unknown

Use Boyle's Law to solve for the unknown pressure (P_2). It is important that the two volumes (V_1 and V_2) are expressed in the same units, so V_2 has been converted to mL.

Step 2: Solve.

First, rearrange the equation algebraically to solve for P_2.

$$P_2 = \frac{P_1 \times V_1}{V_2}$$

Now substitute the known quantities into the equation and solve.

$$P_2 = \frac{387 \text{ kPa} \times 425 \text{ mL}}{1750 \text{ mL}} = 94.0 \text{ kPa}$$

Step 3: Think about your result.

The volume has increased to slightly over 4 times its original value and so the pressure is decreased by about one fourth. The pressure is in kPa and the value has three significant figures. Note that any pressure or volume units can be used as long as they are consistent throughout the problem.

Summary

1.8.4: Charles's Law- Volume and Temperature

Learning Objectives

- Learn and apply Charles's Law.

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stop of molecular motion.

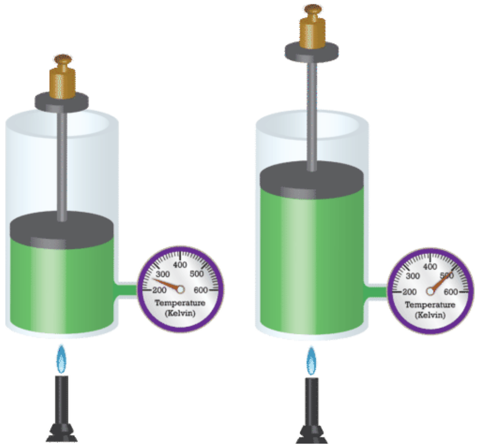


Figure 1.8.4.1: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 1.8.4.1: Temperature-Volume Data

Temperature (K)	**Volume** (mL)	$\boldsymbol{\frac{V}{T} = k} \left(\frac{\text{mL}}{\text{K}}\right)$
50	20	0.40
100	40	0.40
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

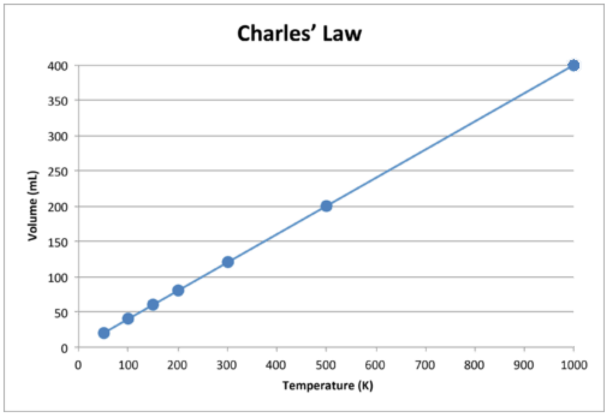


Figure 1.8.4.2:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\text{K} = {}^\circ\text{C} + 273$.

✓ Example 1.8.4.1:

A balloon is filled to a volume of 2.20 L at a temperature of 22°C. The balloon is then heated to a temperature of 71°C. Find the new volume of the balloon.

Solution

Solutions to Example 11.5.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 2.20\text{ L}$ and $T_1 = 22^\circ\text{C} = 295\text{ K}$ $T_2 = 71^\circ\text{C} = 344\text{ K}$ Find: V_2 = ? L
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times T_2}{T_1}$$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{2.20\text{ L} \times 344\ \cancel{\text{K}}}{295\ \cancel{\text{K}}} = 2.57\text{ L}$$
Think about your result.	The volume increases as the temperature increases. The result has three significant figures.

Exercise 1.8.4.1

If V_1 = 3.77 L and T_1 = 255 K, what is V_2 if T_2 = 123 K?

Answer

1.82 L

Example 1.8.4.2:

A sample of a gas has an initial volume of 34.8 L and an initial temperature of −67°C. What must be the temperature of the gas for its volume to be 25.0 L?

Solution

Solutions to Example 11.5.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Given:T_1 = -27^{o}C and V_1 = 34.8 L V_2 = 25.0 L Find: T_2 = ? K
List other known quantities.	K = -27^{o}C + 273
Plan the problem.	1. Convert the initial temperature to Kelvin 2. Rearrange the equation algebraically to solve for T_2. $T_2 = \frac{V_2 \times T_1}{V_1}$
Cancel units and calculate.	1. −67°C + 273 = 206 K 2. Substitute the known quantities into the equation and solve. $T_2 = \frac{25.0 \ \cancel{\text{L}} \times 206 \text{ K}}{34.8 \ \cancel{\text{L}}} = 148 \text{ K}$
Think about your result.	This is also equal to −125°C. As temperature decreases, volume decreases—which it does in this example.

Exercise 1.8.4.2

If V_1 = 623 mL, T_1 = 255°C, and V_2 = 277 mL, what is T_2?

Answer

235 K, or −38°C

Summary

- Charles's Law relates the volume and temperature of a gas at constant pressure and amount.

1.8.5: Combined Gas Law

The modern refrigerator takes advantage of the gas laws to remove heat from a system. Compressed gas in the coils is allowed to expand. This expansion lowers the temperature of the gas and transfers heat energy from the material in the refrigerator to the gas. As the gas is pumped through the coils, the pressure on the gas compresses it and raises the gas temperature. This heat is then dissipated through the coils into the outside air. As the compressed gas is pumped through the system again, the process repeats itself.

Combined Gas Law

To this point, we have examined the relationships between any two of the variables of P, V, and T, while the third variable is held constant. However, situations do arise where all three variables change. The **combined gas law** expresses the relationship between the pressure, volume, and absolute temperature of a fixed amount of gas. For a combined gas law problem, only the amount of gas is held constant.

$$\frac{P \times V}{T} = k \text{ and } \frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

Example 1.8.5.1

2.00 L of a gas at 35^oC and 0.833 atm is brought to standard temperature and pressure (STP). What will be the new gas volume?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P_1 = 0.833 \text{ atm}$
- $V_1 = 2.00 \text{ L}$
- $T_1 = 35^\text{o}\text{C} = 308 \text{ K}$
- $P_2 = 1.00 \text{ atm}$
- $T_2 = 0^\text{o}\text{C} = 273 \text{ K}$

Unknown

Use the combined gas law to solve for the unknown volume $\left(V_2 \right)$. STP is 273 K and 1 atm. The temperatures have been converted to Kelvin.

Step 2: Solve.

First, rearrange the equation algebraically to solve for V_2.

$$V_2 = \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1}$$

Now substitute the known quantities into the equation and solve.

$$V_2 = \frac{0.833 \text{ atm} \times 2.00 \text{ L} \times 273 \text{ K}}{1.00 \text{ atm} \times 308 \text{ K}} = 1.48 \text{ L}$$

Step 3: Think about your result.

Both the increase in pressure and the decrease in temperature cause the volume of the gas sample to decrease. Since both changes are relatively small, the volume does not decrease dramatically.

It may seem challenging to remember all the different gas laws introduced so far. Fortunately, Boyle's, Charles's, and Gay-Lussac's laws can all be easily derived from the combined gas law. For example, consider a situation where a change occurs in the volume and pressure of a gas while the temperature is being held constant. In that case, it can be said that $T_1 = T_2$. Look at the combined gas law and cancel the T variable out from both sides of the equation. What is left over is Boyle's Law: $P_1 \times V_1 = P_2 \times V_2$. Likewise, if the pressure is constant, then $P_1 = P_2$ and cancelling P out of the equation leaves Charles's Law. If the volume is constant, then $V_1 = V_2$ and cancelling V out of the equation leaves Gay-Lussac's Law.

Summary

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

1.8.5.1: Practice Combined Gas Law

This also has practice with Boyle's Law, Charles' Law, and Gay-Lussac's Law, which can be derived from the combined gas law:

Combined Gas Law	At conditions of	Becomes	Also Known As
$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	Constant Temperature ($T_1 = T_2$)	$P_1V_1 = P_2V_2$	Boyle's Law
$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	Constant Pressure ($P_1 = P_2$)	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	Charles' Law
$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	Constant Volume ($V_1 = V_2$)	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	Gay-Lussac's Law

Exercise 1.8.5.1.1

You have a sample of gas with a pressure of 1.86 atm, volume of 4.33 L, and temperature of 26.5 °C. If you cool it to 12.7 °C and decrease the volume to 3.45 L, what will the pressure be?.

Answer

2.23 atm.

Exercise 1.8.5.1.1

You have a flexible container of gas with a volume of 5.220 L at 19.4 °C. At what temperature would the volume increase to be 6.000 L?

Answer

336.3 K or 63.1 °C.

1.8.6: Avogadro’s Law- Volume and Moles

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. How much air should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles (or number of particles) of gas when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

where n is the number of moles of gas and k is a constant. Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

✓ Example 1.8.6.1

A balloon has been filled to a volume of $1.90 \text{ L}$with $0.0920 \text{ mol}$of helium gas. If $0.0210 \text{ mol}$of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

Solution

Solutions to Example 11.11.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 1.90 \text{ L}$ $n_1 = 0.0920 \text{ mol}$ Find: $V_2 = ? \text{ L}$
List other known quantities.	Note that the final number of moles has to be calculated by adding the original number of moles to the moles of added helium. $n_2 = 0.0920 + 0.0210 = 0.1130 \text{ mol}$
Plan the problem.	First, rearrange the equation algebraically to solve for V_2. $$V_2 = \frac{V_1 \times n_2}{n_1}$$
Calculate.	Now substitute the known quantities into the equation and solve. $$V_2 = \frac{1.90 \text{ L} \times 0.1130 \ \cancel{\text{mol}}}{0.0920 \ \cancel{\text{mol}}} = 2.33 \text{ L}$$
Think about your result.	Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly.

? Exercise 1.8.6.1

A 12.8 L volume of gas contains .000498 moles of oxygen gas. At constant temperature and pressure, what volume does .0000136 moles of the gas fill?

Answer

0.350 L

Summary

- Calculations for relationships between volume and number of moles of a gas can be performed using Avogadro's Law.

1.8.7: Ideal Gas Law

There are a number of chemical reactions that require ammonia. In order to carry out a reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

Ideal Gas Law

The combined gas law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. Putting these laws together gives us the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **ideal gas law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable R for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R$$

The ideal gas law is conveniently rearranged to look this way, with the multiplication signs omitted:

$$PV = nRT$$

The variable R in the equation is called the **ideal gas constant**.

Evaluating the Ideal Gas Constant

The value of R, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: kPa, atm, or mm Hg. Therefore, R can have three different values.

We will demonstrate how R is calculated when the pressure is measured in kPa. Recall that the volume of 1.00 mol of any gas at STP is measured to be 22.414 L We can substitute 101.325 kPa for pressure, 22.414 L for volume, and 273.15 K for temperature into the ideal gas equation and solve for R.

$$R = \frac{PV}{nT} = \frac{101.325 \text{ kPa} \times 22.414 \text{ L}}{1.000 \text{ mol} \times 273.15 \text{ K}} = 8.314 \text{ kPa} \cdot \text{L/K} \cdot \text{mol}$$

This is the value of R that is to be used in the ideal gas equation when the pressure is given in kPa. The table below shows a summary of this and the other possible values of R. It is important to choose the correct value of R to use for a given problem.

Table 1.8.7.1: Values of the Ideal Gas Constant

Unit of P	Unit of V	Unit of n	Unit of T	Value and Unit of R
kPa	L	mol	K	$8.314 \text{ J/K} \cdot \text{mol}$
atm	L	mol	K	$0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}$
mm Hg	L	mol	K	$62.36 \text{ L} \cdot \text{mm Hg/K} \cdot \text{mol}$

Notice that the unit for R when the pressure is in kPa has been changed to $\text{J/K} \cdot \text{mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

Example 1.8.7.1

What volume is occupied by $3.760\ \mathrm{g}$ of oxygen gas at a pressure of $88.4\ \mathrm{kPa}$ and a temperature of $19^\mathrm{o}\mathrm{C}$? Assume the oxygen is an ideal gas.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P = 88.4\ \mathrm{kPa}$
- $T = 19^\mathrm{o}\mathrm{C} = 292\ \mathrm{K}$
- $\text{Mass } \mathrm{O_2} = 3.760\ \mathrm{g}$
- $\mathrm{O_2} = 32.00\ \mathrm{g/mol}$
- $R = 8.314\ \mathrm{J/K \cdot mol}$

Unknown

In order to use the ideal gas law, the number of moles of $\mathrm{O_2}$ (n) must be found from the given mass and the molar mass. Then, use $PV = nRT$ to solve for the volume of oxygen.

Step 2: Solve.

$$3.760\ \mathrm{g} \times \frac{1\ \mathrm{mol\ O_2}}{32.00\ \mathrm{g\ O_2}} = 0.1175\ \mathrm{mol\ O_2}$$

Rearrange the ideal gas law and solve for V.

$$V = \frac{nRT}{P} = \frac{0.1175\ \mathrm{mol} \times 8.314\ \mathrm{J/K \cdot mol} \times 292\ \mathrm{K}}{88.4\ \mathrm{kPa}} = 3.23\ \mathrm{L\ O_2}$$

Step 3: Think about your result.

The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar volume $(22.4\ \mathrm{L/mol})$ since the pressure and temperature are reasonably close to standard. The result has three significant figures because of the values for T and P. Since a joule $(\mathrm{J}) = \mathrm{kPa} \cdot \mathrm{L}$, the units cancel out correctly, leaving a volume in liters.

Summary

- The ideal gas constant is calculated.
- An example of calculations using the ideal gas law is shown.

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1.8.7.1: Practice Ideal Gas Law

Exercise 1.8.7.1.1

How many moles of gas would be in a 3.75 L container of gas that has a pressure of 2.44 atm and a temperature of 303.0 K?

Answer

0.368 moles.

Exercise 1.8.7.1.1

At what temperature would 0.850 moles of gas have a volume of 11.0 L and a pressure of 855 mm Hg? (Remember 1 atm = 760 mm Hg exactly)

Answer

177 K or –96 °C.

Exercise 1.8.7.1.1

What volume would 2.33 moles of gas take up at a temperature of 313 K and a pressure of 1.38 atm?

Answer

43.4 L.

1.8.8: Mixtures of Gases - Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen

Learning Objectives

- Explain Dalton's Law of Partial Pressures.

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are 96.5% carbon dioxide and 3% nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus would contribute a pressure well over $2700 \: \text{mm Hg}$. And there is no oxygen present, so we couldn't breathe there. Not that we would want to go to Venus, as the surface temperature is usually over 460^oC.

Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about 78% nitrogen and 21% oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up 78% of the gas particles in a given sample of air, it exerts 78% of the pressure. If the overall atmospheric pressure is $1.00 \: \text{atm}$, then the pressure of just the nitrogen in the air is $0.78 \: \text{atm}$. The pressure of the oxygen in the air is $0.21 \: \text{atm}$.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of nitrogen is represented by P_{N_2}. **Dalton's Law of Partial Pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's Law can be expressed with the following equation:

$$P_\text{total} = P_1 + P_2 + P_3 + \cdots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, P_1 and P_2, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300 \: \text{mm Hg}$ and $P_2 = 500 \: \text{mm Hg}$, then $P_\text{total} = 800 \: \text{mm Hg}$.

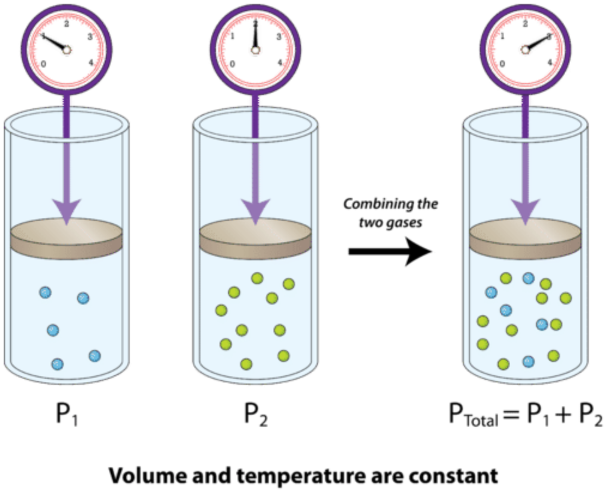

Figure $\PageIndex{1}$: Dalton's Law states that the pressure of a gas mixture is equal to the partial pressures of the combining gases.

Collecting Gases Over Water

You need to do a lab experiment where hydrogen gas is generated. In order to calculate the yield of gas, you have to know the pressure inside the tube where the gas is collected. But how can you get a barometer in there? Very simple: you don't. All you need is the atmospheric pressure in the room. As the gas pushes out the water, it is pushing against the atmosphere, so the pressure inside is equal to the pressure outside.

Gas Collection by Water Displacement

Gases that are produced in laboratory experiments are often collected by a technique called **water displacement** (Figure $1.8.8.2$). A bottle is filled with water and placed upside-down in a pan of water. The reaction flask is fitted with rubber tubing, which is then fed under the bottle of water. As the gas is produced in the reaction flask, it exits through the rubber tubing and displaces the water in the bottle. When the bottle is full of the gas, it can be sealed with a lid.

Figure $1.8.8.2$: A gas produced in a chemical reaction can be collected by water displacement.

Because the gas is collected over water, it is not pure, but is mixed with vapor from the evaporation of the water. Dalton's Law can be used to calculate the amount of the desired gas by subtracting the contribution of the water vapor.

$$P_{\text{Total}} = P_g + P_{H_2O}$$

where P_g is the pressure of the desired gas, which can be solved for:

$$P_g = P_{Total} - P_{H_2O}$$

In order to solve a problem, it is necessary to know the vapor pressure of water at the temperature of the reaction (see table below). The sample problem illustrates the use of Dalton's Law when a gas is collected over water.

Table $1.8.8.1$: Vapor Pressure of Water (mm Hg) at Selected Temperatures $(^\circ\text{C})$

0	**5**	**10**	**15**	**20**	**25**	**30**	**35**	**40**	**45**	**50**	**55**	**60**
4.58	6.54	9.21	12.79	17.54	23.76	31.82	42.18	55.32	71.88	92.51	118.04	149.38

✓ Example 14.14.1

A certain experiment generates $2.58\ \text{L}$ of hydrogen gas, which is collected over water. The temperature is 20°C and the atmospheric pressure is $98.60\ \text{kPa}$ Find the volume that the dry hydrogen would occupy at STP.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $V_{\text{Total}} = 2.58\ \text{L}$
- $T = 20^\circ\text{C} = 293\ \text{K}$
- $P_{\text{Total}} = 98.60\ \text{kPa} = 739.7\ \text{mm Hg}$

Unknown

- V_{H_2} at STP $=?\ \text{L}$

The atmospheric pressure is converted from kPa to mm Hg in order to match units with the table. The sum of the pressures of the hydrogen and the water vapor is equal to the atmospheric pressure. The pressure of the hydrogen is found by subtraction. Then, the volume of the gas at STP can be calculated by using the combined gas law.

Step 2: Solve.

$$\begin{aligned} P_{H_2} &= P_{\text{Total}} - P_{H_2O} \\ &= 739,7\ \text{mm Hg} - 17.54\ \text{mm Hg} \\ &= 722.2\ \text{mm Hg} \end{aligned}$$

Now the combined gas law is used, solving for V_2, the volume of hydrogen at STP.

$$\begin{aligned} V_2 &= \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1} \\ &= \frac{722.2 \text{ mm Hg} \times 2.58 \text{ L} \times 273 \text{ K}}{760 \text{ mm Hg} \times 293 \text{ K}} \\ &= 2.28 \text{ L H}_2 \end{aligned}$$

Step 3: Think about your result.

If the hydrogen gas were to be collected at STP and without the presence of the water vapor, its volume would be 2.28 L. This is less than the actual collected volume because some of that is water vapor. The conversion using STP is useful for stoichiometry purposes.

Summary

- Dalton's Law of Partial Pressures states that the total pressure in a system is equal to the sum of the partial pressures of the gases present.
- The vapor pressure due to water in a sample can be corrected for, in order to get the true value for the pressure of the gas.

1.8.8.1: Practice Dalton's Law

Exercise 1.8.8.1.1

You have a tank with a mixture of nitrogen and oxygen gas. The total pressure of the tank is 2550 psi. If the pressure of nitrogen is 1270 psi, what is the pressure of the oxygen?

Answer

1280 psi.

Exercise 1.8.8.1.1

A sample of hydrogen gas is collected over water in a flask that has the same pressure as the lab room (0.9865 atm that day). The flask of water is 22.0 °C, and at that temperature, water vapor pressure is 19.8 mm Hg. What is the pressure of the hydrogen gas collected? If it occupies a volume of 445 mL, how many moles of hydrogen gas are there? (Assume the temperature of the hydrogen gas is also 22.0 °C)

Answer

Pressure = 0.9604 atm (or 729.9 mm Hg). There are 0.0177 moles.

1.8.9: Conversions Between Moles and Gas Volume

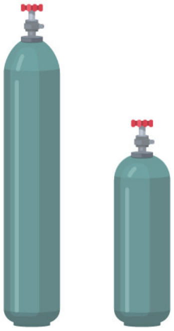

Figure 1.8.9.1 (Credit: CK-12 Foundation; Source: CK-12 Foundation; License: CK-12 Curriculum Materials license)

How can you tell how much gas is in these containers?

Small gas tanks are often used to supply gases for chemistry reactions. A gas gauge will give some information about how much is in the tank, but quantitative estimates are needed so that the reaction will be able to proceed to completion. Knowing how to calculate the necessary parameters for gases is very helpful to avoid running out earlier than desired.

Conversions Between Moles and Gas Volume

Molar volume at STP can be used to convert from moles to gas volume and from gas volume to moles. The equality of $1 \text{ mol} = 22.4 \text{ L}$ is the basis for the conversion factor.

Example 1.8.9.1: Converting Gas Volume to Moles

Many metals react with acids to produce hydrogen gas. A certain reaction produces 86.5 L of hydrogen gas at STP. How many moles of hydrogen were produced?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 86.5 L H_2
- $1 \text{ mol} = 22.4 \text{ L}$

Unknown

- moles of H_2

Apply a conversion factor to convert from liters to moles.

Step 2: Calculate.

$$86.5 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} = 3.86 \text{ mol H}_2$$

Step 3: Think about your result.

The volume of gas produced is nearly four times larger than the molar volume. The fact that the gas is hydrogen plays no role in the calculation.

Example 1.8.9.2: Convertig Moles to Gas Volume

What volume does 4.96 mol of O_2 occupy at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $4.96 \text{ mol } O_2$
- $1 \text{ mol} = 22.4 \text{ L}$

Unknown

- volume of O_2

Step 2: Calculate.

$$4.96 \text{ mol} \times 22.4 \text{ L/mol} = 111.1 \text{ L}$$

Step 3: Think about your result.

The volume seems correct given the number of moles.

Example 1.8.9.3: Converting Volume to Mass

If we know the volume of a gas sample at STP, we can determine how much mass is present. Assume we have 867 L of N_2 at STP. What is the mass of the nitrogen gas?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $867 \text{ L } N_2$
- $1 \text{ mol} = 22.4 \text{ L}$
- Molar mass $N_2 = 28.02 \text{ g/mol}$

Unknown

- mass of N_2

Step 2: Calculate.

We start by determining the number of moles of gas present. We know that 22.4 liters of a gas at STP equals one mole, so:

$$867 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 38.7 \text{ mol}$$

We also know the molecular weight of N_2 (28.0 g/mol) so we can then calculate the weight of nitrogen gas in 867 liters:

$$38.7 \text{ mol} \times \frac{28 \text{ g}}{1 \text{ mol}} = 1083.6 \text{ g } N_2$$

Step 3: Think about your result.

In a multi-step problem, be sure that the units check out.

Summary

- Conversions between moles and volume of a gas are shown.

Review

1. In the problems above, why was the gas always at standard temperature and pressure?
2. A container contains 45.2 L of N_2 gas at STP. How many moles of N_2 gas are in the container?
3. If the gas in the previous problem was CH_4 gas at STP instead of N_2 gas, then how many moles of CH_4 gas would there be?

1.8.9.1: Practice Gas Stoichiometry

Exercise 1.8.9.1.1

If you reacted 14.7 g of Zn in the following reaction, what volume of H_2 gas, measured at STP, would be produced?

$Zn_{(s)} + 2\ HCl_{(aq)}$ --> $ZnCl_{2\ (aq)} + H_{2\ (g)}$

Answer

5.04 L. (0.225 moles)

Exercise 1.8.9.1.1

What mass of O_2 gas is contained in a 3.85 L container at STP?.

Answer

5.50 g. (0.172 moles)

Exercise 1.8.9.1.1

How many liters of H_2 gas, measured at STP, is needed to produce 85.0 g of NH_3?

$3\ H_{2\ (g)} + N_{2\ (g)}$ --> $2\ NH_{3\ (g)}$

Answer

168 L. (7.48 moles)

SECTION OVERVIEW

1.9: Solutions

1.9.1: Solutions - Homogeneous Mixtures

Learning Objectives

- Learn terminology involving solutions.
- Explain the significance of the statement "like dissolves like."
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater or lesser presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious.

Figure 1.9.1.1: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent. Table 1.9.1.1lists some common types of solutions, with examples of each.

Table 1.9.1.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	salt water
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

✓ Example 1.9.1.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose $(C_{12}H_{22}O_{11})$ in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 1.9.1.1

A solution is made by dissolving 3.33 g of $\mathrm{HCl(g)}$ in 40.0 g of liquid methyl alcohol $(\mathrm{CH_3OH})$. Identify the solvent and solute in the resulting solution.

Answer

solute: $\mathrm{HCl(g)}$

solvent: $\mathrm{CH_3OH}$

Like Dissolves Like

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, does dissolve in water—it is this oxygen that the fish take in through their gills. The reason we can enjoy carbonated sodas is also due to a nonpolar compound that dissolves in water. Pepsi-cola and all the other sodas have carbon dioxide gas, $\mathrm{CO_2}$, a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Table 1.9.1.2: Summary of Solubilities

Solute (Polarity of Compound)	Solvent (Polarity of Compound)	Dominant Intermolecular Force	Is Solution Formed?
Polar	Polar	Dipole-Dipole Force and/or Hydrogen Bond	yes
Non-polar	Non-polar	Dispersion Force	yes
Polar	Non-polar		no
Non-polar	Polar		no
Ionic	Polar	Ion-Dipole	yes
Ionic	Non-polar		no

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges, a solution does not form. When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in Chapter 10). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water.

Figure 1.9.1.2: Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

✓ Example 1.9.1.2: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O? Explain your answer.

Solution

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

? Exercise 1.9.1.2

Would C_3H_7OH be more soluble in CCl_4 or H_2O? Explain your answer.

Answer

H_2O, because both experience hydrogen bonding.

✓ Example 1.9.1.3

Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol (CH_3OH)
b. sodium sulfate (\ce{Na2SO4}\))
c. octane (C_8H_{18})

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

a. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
b. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
c. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

? Exercise 1.9.1.3: Toluene

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer

Octane (C_8H_{18}) will dissolve. It is also non-polar.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.

1.9.2: Solutions of Solids Dissolved in Water- How to Make Rock Candy

Learning Objectives

- Define electrolytes and non electrolytes
- Explain why solutions form.
- Discuss the idea of water as the "universal solvent".
- Explain how water molecules attract ionic solids when they dissolve in water.

We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. We also know that solutions have constant composition, and that this composition can be varied up to a point to maintain the homogeneous nature of the solution. But how exactly do solutions form? Why is it that oil and water will not form a solution, and yet vinegar and water will? Why could we dissolve table salt in water, but not in vegetable oil? The reasons why solutions will form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Solubility and Saturation

Table salt (NaCl) readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is specified as the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table $1.9.2.1$ lists the solubilities of some simple ionic compounds. These solubilities vary widely. NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table $1.9.2.1$: Solubilities of Some Ionic Compounds

Solute	**Solubility (g per 100 g of H_2O at 25°C)**
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated, but rather concentrated. In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

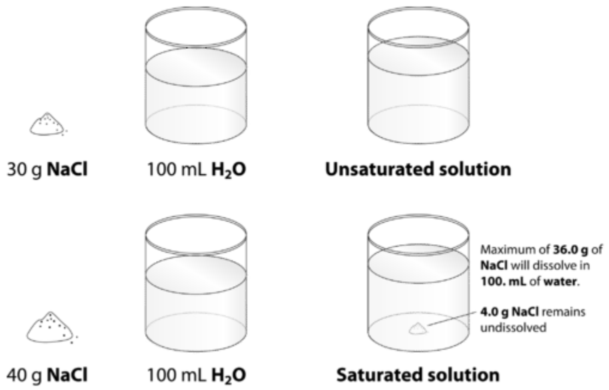

Figure 1.9.2.1: When 30.0 g of NaCl is added to 100 mL, it all dissolves, forming an unsaturated solution. When 40.0 g is added, 36.0 g dissolves and 4.0 g remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

Electrolyte Solutions: Dissolved Ionic Solids

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte (good conductor). If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, the substance is a weak electrolyte (does not conduct electricity as well).

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 1.9.2.1).

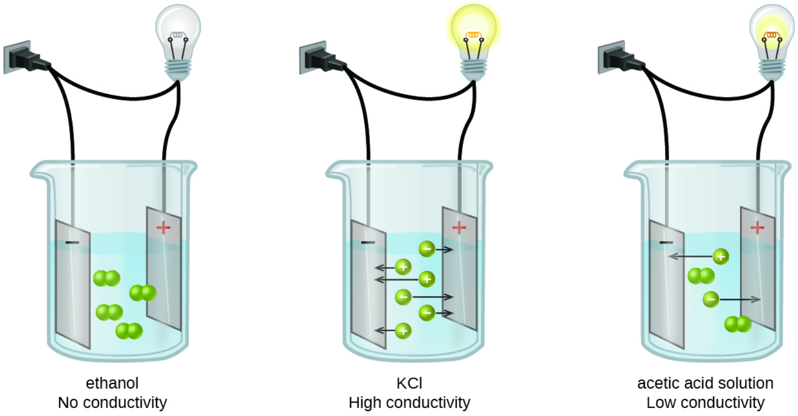

Figure 1.9.2.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte. This diagram shows three separate beakers. Each has a wire plugged into a wall outlet. In each case, the wire leads from the wall to the beaker and is split resulting in two ends. One end leads to a light bulb and continues on to a rectangle labeled with a plus sign. The other end leads to a rectangle labeled with a minus sign. The rectangles are in a solution. In the first beaker, labeled “Ethanol No Conductivity,” four pairs of linked small green spheres suspended in the solution between the rectangles. In the second beaker, labeled “K C l Strong Conductivity,” six individual green spheres, three labeled plus and three labeled minus are suspended in the solution. Each of the six spheres has an arrow extending from it pointing to the rectangle labeled with the opposite sign. In the third beaker, labeled “Acetic acid solution Weak conductivity,” two pairs of joined green spheres and two individual spheres, one labeled plus and one labeled minus are shown suspended between the two rectangles. The plus labeled sphere has an arrow pointing to the rectangle labeled minus and the minus labeled sphere has an arrow pointing to the rectangle labeled plus.

Water and other polar molecules are attracted to ions, as shown in Figure 1.9.2.2 The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

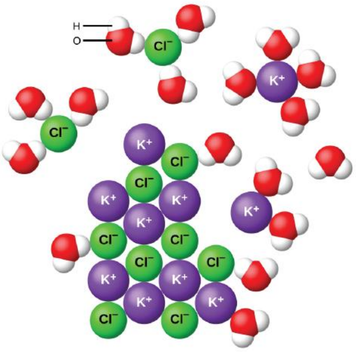

Figure 1.9.2.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown. The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled C l superscript minus mixed and touching near the center of the diagram. Outside of this cluster of spheres are seventeen clusters of three spheres, which include one red and two white spheres. A red sphere in one of these clusters is labeled O. A white sphere is labeled H. Two of the green C l superscript minus spheres are surrounded by three of the red and white clusters, with the red spheres closer to the green spheres than the white spheres. One of the K superscript plus purple spheres is surrounded by four of the red and white clusters. The white spheres of these clusters are closest to the purple spheres.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

✓ Example 1.9.2.1: Identifying Ionic Compounds

Which compound(s) will dissolve in solution to separate into ions?

a. LiF

b. P_2F_5
c. C_2H_5OH

Solution

LiF will separate into ions when dissolved in solution, because it is an ionic compound. P_2F_5 and C_2H_5OH are both covalent and will stay as molecules in a solution.

? Exercise 1.9.2.1

Which compounds will dissolve in solution to separate into ions?

a. $C_6H_{12}O_{11}$, glucose
b. CCl_4
c. $CaCl_2$
d. $AgNO_3$

Answer

c & d

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per 100 g of solvent. The solubility of sodium chloride in water is 36.0 g per 100 g water at $20^{\circ}C$. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (Figure 1.9.2.4).

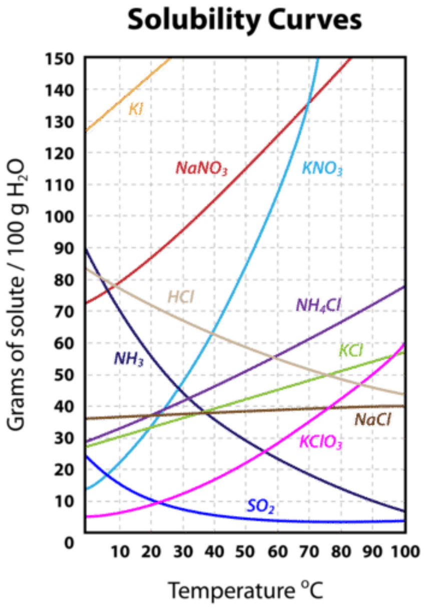

Figure 1.9.2.4: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, $\mathrm{NH_3}$, and $\mathrm{SO_2}$—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that 80 g of $\mathrm{KNO_3}$ is added to 100 g of water at $30^\mathrm{o}\mathrm{C}$. According to the solubility curve, approximately 48 g of $\mathrm{KNO_3}$ will dissolve at $30^\mathrm{o}\mathrm{C}$. This means that the solution will be saturated since 48 g is less than 80 g. We can also determine that there will be $80 - 48 = 32$ g of undissolved $\mathrm{KNO_3}$ remaining at the bottom of the container. Now suppose that this saturated solution is heated to $60^\mathrm{o}\mathrm{C}$. According to the curve, the solubility of $\mathrm{KNO_3}$ at $60^\mathrm{o}\mathrm{C}$ is about 107 g. Now the solution is unsaturated since it contains only the original 80 g of dissolved solute. Now suppose the solution is cooled all the way down to $0^\mathrm{o}\mathrm{C}$. The solubility at $0^\mathrm{o}\mathrm{C}$ is about 14 g, meaning that $80 - 14 = 66$ g of the $\mathrm{KNO_3}$ will re-crystallize.

Summary

- Solubility is the specific amount of solute that can dissolve in a given amount of solvent.
- Saturated and unsaturated solutions are defined.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).
- Water is considered the universal solvent since it can dissolve both ionic and polar solutes, as well as some nonpolar solutes (in very limited amounts).
- The solubility of a solid in water increases with an increase in temperature.

Vocabulary

- **Miscible** - Liquids that have the ability to dissolve in each other.
- **Immiscible** - Liquids that do not have the ability to dissolve in each other.
- **Electrostatic attraction** - The attraction of oppositely charged particles.

1.9.3: Henry's Law

Having a soft drink in outer space poses some special problems. Under microgravity, the carbonation can quickly dissipate if not kept under pressure. You can't open the can, or you will lose carbonation. So, a special pressurized container has been developed to get around the problem of gas loss at low gravity.

Henry's Law

Pressure has very little effect on the solubility of solids or liquids, but has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the $\mathrm{CO_2}$ molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more $\mathrm{CO_2}$ is added to the space above the container, causing a pressure increase. In this case, more $\mathrm{CO_2}$ molecules are in contact with the water and so more of them dissolve. Thus, the solubility increases as the pressure increases. As with a solid, the $\mathrm{CO_2}$ that is undissolved reaches an equilibrium with the dissolved $\mathrm{CO_2}$, represented by the equation:

$$\mathrm{CO_2}\,(g) \rightleftharpoons \mathrm{CO_2}\,(aq)$$

At equilibrium, the rate of gaseous $\mathrm{CO_2}$ dissolution is equal to the rate of dissolved $\mathrm{CO_2}$ coming out of the solution.

When carbonated beverages are packaged, they are done so under high $\mathrm{CO_2}$ pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is open, the equilibrium is disrupted because the $\mathrm{CO_2}$ pressure above the liquid decreases. Immediately, bubbles of $\mathrm{CO_2}$ rapidly exit the solution and escape out of the top of the open bottle. The amount of dissolved $\mathrm{CO_2}$ decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more $\mathrm{CO_2}$ comes out of the liquid.

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774-1836). **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. Henry's law can be written as follows:

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 and P_1 are the solubility and the pressure at an initial set of conditions; S_2 and P_2 are the solubility and pressure at another changed set of conditions. The solubility of a gas is typically reported in $\mathrm{g/L}$.

Example 1.9.3.1

The solubility of a certain gas in water is $0.745\ \mathrm{g/L}$ at standard pressure. What is its solubility when the pressure above the solution is raised to $4.50\ \mathrm{atm}$? The temperature is constant at $20^\circ\mathrm{C}$.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $S_1 = 0.745\ \mathrm{g/L}$
- $P_1 = 1.00\ \mathrm{atm}$
- $P_2 = 4.50\ \mathrm{atm}$

Unknown

Substitute into Henry's law and solve for S_2.

Step 2: Solve.

$$S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.745\ \mathrm{g/L} \times 4.50\ \mathrm{atm}}{1.00\ \mathrm{atm}} = 3.35\ \mathrm{g/L}$$

Step 3: Think about your result.

The solubility is increased to 4.5 times its original value, according to the direct relationship.

Summary

1.9.3.1: Practice Solution Properties

Exercise 1.9.3.1.1

Which of the following would conduct electricity well if dissolved in water?

A) glucose ($C_6H_{12}O_6$)

B) baking soda ($NaHCO_3$)

C) ethanol (C_2H_6O)

Answer

B) baking soda (sodium bicarbonate) is ionic and dissociates into ions when dissolved.

Exercise 1.9.3.1.1

Which of the following would result in increased solubility of a gas in water?

A) increased pressure

B) increased temperature

Answer

A) increased pressure, according to Henry's Law, increases the solubility of a gas. Increased temperature generally decreases the solubility of a gas in water.

1.9.4: Percent Solutions

There are human cultures that do not recognize numbers above three. Anything greater than that is simply referred to as "much" or "many". Although this form of calculation may seem very limited, American culture does the same thing to a certain degree. For example, there are several ways to express the amount of solute in a solution in a quantitative manner. The **concentration** of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution. A **concentrated** solution is one that has a relatively large amount of dissolved solute. A **dilute** solution is one that has a relatively small amount of dissolved solute. However, those terms are vague, and it is often necessary to express concentration with numbers.

Percent Solutions

One way to describe the concentration of a solution is by the percent of a solute in the solvent. The percent can further be determined in one of two ways: (1) the ratio of the mass of the solute divided by the mass of the solution or (2) the ratio of the volume of the solute divided by the volume of the solution.

Mass Percent

When the solute in a solution is a solid, a convenient way to express the concentration is by mass percent $\left(\frac{\text{mass}}{\text{mass}}\right)$, which is the grams of solute per 100 g of solution.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Suppose that a solution was prepared by dissolving 25.0 g of sugar into 100 g of water. The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25\text{ g sugar}}{125\text{ g solution}} \times 100\% = 20\%\text{ sugar}$$

Sometimes you may want to make up a particular mass of solution of a given percent by mass, and need to calculate what mass of the solvent to use. For example, you need to make 3000 g of a 5% solution of sodium chloride. You can rearrange and solve for the mass of solute:

$$\text{mass of solute} = \frac{\text{percent by mass}}{100\%} \times \text{mass of solution} = \frac{5\%}{100\%} \times 3000\text{ g} = 150\text{ g NaCl}$$

You would need to weigh out 150 g of NaCl and add it to 2850 g of water. Notice that it was necessary to subtract the mass of the NaCl (150 g) from the mass of solution (3000 g) to calculate the mass of the water that would need to be added.

Volume Percent

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution, expressed as a percent, yields the percent by volume $\left(\frac{\text{volume}}{\text{volume}}\right)$ of the solution. If a solution is made by adding 40 mL of ethanol to 200 mL of water, the percent by volume is:

$$\begin{aligned}\text{Percent by volume} &= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \\ &= \frac{40\text{ mL ethanol}}{240\text{ mL solution}} \times 100\% \\ &= 16.7\%\text{ ethanol}\end{aligned}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

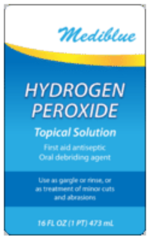

Figure 1.9.4.1: Hydrogen peroxide is commonly labeled as a 3% by volume solution for use as a disinfectant.

Summary

- The concentration of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution.
- A concentrated solution has a relatively large amount of dissolved solute.
- A dilute solution has a relatively small amount of dissolved solute.
- Techniques for calculation of percent mass and percent volume solution concentrations are described.

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1.9.4.1: Practice Mass Percent and Volume Percent

Exercise 1.9.4.1.1

What is the mass percent concentration of a solution made by adding 21.4 g of NaCl to 135.0 g water? (Hint: find mass of solution!)

Answer

13.7 % (m/m)

Exercise 1.9.4.1.1

What is the volume percent of a solution made by adding 38.2 mL of ethanol to enough water to make 250.0 mL of solution?

Answer

15.3 % (v/v)

1.9.5: Molarity

Chemists deal with amounts of molecules every day. Chemical reactions are described as so many molecules of compound A reacting with so many molecules of compound B to form so many molecules of compound C. When we determine how much reagent to use, we need to know the number of molecules in a given volume of the reagent. Percent solutions only tell us the number of grams, not molecules. A 100 mL solution of $2\% \text{ NaCl}$ will have a very different number of molecules than a 2% solution of CsCl. So, we need another way to talk about numbers of molecules.

Molarity

Chemists primarily need the concentration of solutions to be expressed in a way that accounts for the number of particles that react according to a particular chemical equation. Since percentage measurements are based on either mass or volume, they are generally not useful for chemical reactions. A concentration unit based on moles is preferable. The **molarity** (M) of a solution is the number of moles of solute dissolved in one liter of solution. To calculate the molarity of a solution, divide the moles of solute by the volume of the solution expressed in liters:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

Note that the volume is in liters of solution and not liters of solvent. When a molarity is reported, the unit is the symbol M and is read as "molar". For example a solution labeled as 1.5 M NH_3 is read as "1.5 molar ammonia solution".

Example 1.9.5.1

A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Mass} = 42.23 \text{ g NH}_4\text{Cl}$
- $\text{Molar mass NH}_4\text{Cl} = 53.50 \text{ g/mol}$
- $\text{Volume solution} = 500.0 \text{ mL} = 0.5000 \text{ L}$

Unknown

The mass of the ammonium chloride is first converted to moles. Then the molarity is calculated by dividing by liters. Note that the given volume has been converted to liters.

Step 2: Solve.

$$42.23 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.50 \text{ g NH}_4\text{Cl}} = 0.7893 \text{ mol NH}_4\text{Cl}$$

$$\frac{0.7893 \text{ mol NH}_4\text{Cl}}{0.5000 \text{ L}} = 1.579 \text{ M}$$

Step 3: Think about your result.

The molarity is 1.579 M, meaning that a liter of the solution would contain $1.579 \text{ mol NH}_4\text{Cl}$ Four significant figures are appropriate.

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams.

Example 1.9.5.2

A chemist needs to prepare 3.00 L of a 0.250 M solution of potassium permanganate $\left(\text{KMnO}_4\right)$. What mass of KMnO_4 does she need to make the solution?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Molarity} = 0.250 \text{ M}$
- $\text{Volume} = 3.00 \text{ L}$
- $\text{Molar mass KMnO}_4 = 158.04 \text{ g/mol}$

Unknown

Moles of solute is calculated by multiplying molarity by liters. Then, moles is converted to grams.

Step 2: Solve.

$$\text{mol KMnO}_4 = 0.250 \text{ M KMnO}_4 \times 3.00 \text{ L} = 0.750 \text{ mol KMnO}_4$$

$$0.750 \text{ mol KMnO}_4 \times \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} = 119 \text{ g KMnO}_4$$

Step 3: Think about your result.

When 119 g of potassium permanganate is dissolved into water to make 3.00 L of solution, the molarity is 0.250 M.

Summary

- The molarity (M) of a solution is the number of moles of solute dissolved in one liter of solution.
- Calculations using the concept of molarity are described.

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1.9.5.1: Practice Molarity

Exercise 1.9.5.1.1

What molarity of solution is made by dissolving 34.8 g NaCl in enough water to make 300.00 mL of solution?

Answer

1.98 M NaCl solution.

Exercise 1.9.5.1.1

What molarity solution is made by dissolving 25.3 g methanol (CH_4O) in enough water to make 125 mL solution?

Answer

6.32 M methanol.

1.9.6: Dilution

Muriatic acid (another name for HCl) is widely used for cleaning concrete and masonry surfaces. The acid must be diluted before use to get it down to a safer strength. Commercially available at concentrations of about 18%, muriatic acid can be used to remove scales and deposits (usually composed of basic materials).

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the number of moles of the solute does not change, while the volume of the solution increases. We can set up an equality between the moles of the solute before the dilution (1) and the moles of the solute after the dilution (2).

$$\mathrm{mol}_1 = \mathrm{mol}_2$$

Since the moles of solute in a solution is equal to the molarity multiplied by the liters, we can set those equal.

$$M_1 \times L_1 = M_2 \times L_2$$

Finally, because the two sides of the equation are set equal to one another, the volume can be described in any unit that we choose, as long as that unit is the same on both sides. Our equation for calculating the molarity of a diluted solution becomes:

$$M_1 \times V_1 = M_2 \times V_2 \tag{1.9.6.1}$$

Suppose that you have $100.\ \mathrm{mL}$ of a $2.0\ \mathrm{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \mathrm{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2:

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \mathrm{M} \times 100.\ \mathrm{mL}}{500.\ \mathrm{mL}} = 0.40\ \mathrm{M\ HCl}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves deciding how much of a highly concentrated solution is required to make a desired quantity of solution of lesser concentration. The highly concentrated solution is typically referred to as the stock solution.

Example 1.9.6.1

Nitric acid $\left(\mathrm{HNO_3}\right)$ is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \mathrm{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \mathrm{L}$ of a $0.50\ \mathrm{M}$ solution?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Stock $\mathrm{HNO_3} = 16\ \mathrm{M}$
- $V_2 = 8.00\ \mathrm{L}$
- $M_2 = 0.50\ \mathrm{M}$

Unknown

The unknown in the equation is V_1, the volume of the concentrated stock solution.

Step 2: Solve

Using Equation 1.9.6.1 we can solve for V_1:

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.50\ \mathrm{M} \times 8.00\ \mathrm{L}}{16\ \mathrm{M}} = 0.25\ \mathrm{L} = 250\ \mathrm{mL}$$

Step 3: Think about your result

$250\ \mathrm{mL}$ of the stock $\mathrm{HNO_3}$ needs to be diluted with water to a final volume of $8.00\ \mathrm{L}$. The dilution is by a factor of 32 to go from $16\ \mathrm{M}$ to $0.5\ \mathrm{M}$.

Dilutions can be performed in the laboratory with various tools, depending on the volumes required and the desired accuracy. The images below illustrate the use of two different types of pipettes. In the first figure, a glass **pipette** is being used to transfer a portion of a solution to a graduated cylinder. Use of a pipette rather than a graduated cylinder for the transfer improves accuracy. The second figure shows a **micropipette**, which is designed to quickly and accurately dispense small volumes. Micropipettes are adjustable and come in a variety of sizes.

Figure 1.9.6.1: Volumetric pipette and micropipette.

Summary

- A process is described to calculate dilutions.
- Pipettes and micropipettes are used to transfer and dispense solution.

1.9.6.1: Practice Dilutions

Exercise 1.9.6.1.1

If you add water to 35.0 mL of 6.75 M H_2SO_4 solution until you have 250.0 mL of new solution, what is the concentration of sulfuric acid in that new solution?

Answer

0.945 M H_2SO_4.

Exercise 1.9.6.1.1

What volume of 4.00 M NaOH solution would you need to use in order to make 575 mL of a 1.20 M solution?

Answer

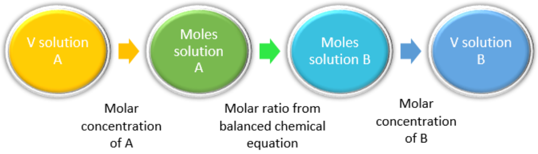
173 mL.

1.9.7: Solution Stoichiometry

Learning Objectives

- Determine amounts of reactants or products in aqueous solutions.

As we learned previously, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are "switched" (they *replace* each other). Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and therefore their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead (II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead (II) chloride.

$$\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow PbCl_2(s) + 2\,NaNO_3(aq)} \tag{1.9.7.1}$$

In the reaction shown above, if we mixed 0.123 L of a 1.00 M solution of NaCl with 1.50 M solution of $\mathrm{Pb(NO_3)_2}$, we could calculate the volume of $\mathrm{Pb(NO_3)_2}$ solution needed to completely precipitate the $\mathrm{Pb^{2+}}$ ions.

The molar concentration can also be expressed as the following:

$$1.00\,M\,\mathrm{NaCl} = \frac{1.00\,mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}}$$

and

$$1.50\,M\,\mathrm{Pb(NO_3)_2} = \frac{1.50\,mol\ \mathrm{Pb(NO_3)_2}}{1\ L\ \mathrm{Pb(NO_3)_2 solution}}$$

First, we must examine the reaction stoichiometry in the balanced reaction (Equation 1.9.7.1). In this reaction, one mole of $\mathrm{Pb(NO_3)_2}$ reacts with two moles of NaCl to give one mole of $\mathrm{PbCl_2}$ precipitate. Thus, the concept map utilizing the stoichiometric ratios is:

so the volume of lead (II) nitrate that reacted is calculated as:

$$0.123\ L\ \mathrm{NaCl\,solution} \times \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}} \times \frac{1\ mol\ \mathrm{Pb(NO_3)_2}}{2\ mol\ \mathrm{NaCl}} \times \frac{1\ L\ \mathrm{Pb(NO_3)_2\ solution}}{1.5\ mol\ \mathrm{Pb(NO_3)_2}} = 0.041\ \mathrm{Pb(NO_3)_2}$$
$$L\ \mathrm{solution}$$

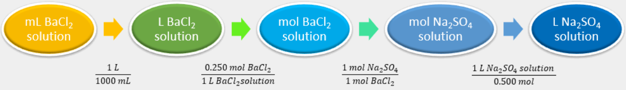
This volume makes intuitive sense for two reasons: (1) the number of moles of $\mathrm{Pb(NO_3)_2}$ required is half of the number of moles of NaCl, based off of the stoichiometry in the balanced reaction (Equation 1.9.7.1); (2) the concentration of $\mathrm{Pb(NO_3)_2}$ solution is 50% greater than the NaCl solution, so less volume is needed.

Example 1.9.7.1

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate all $\mathrm{Ba^{2+}}$ in the solution?

Solution

Solutions to Example 13.8.1

Steps for Problem Solving	Example 1.9.7.1
Identify the "given" information and what the problem is asking you to "find."	Given: 275 mL $\mathrm{BaCl_2}$ 0.250 M $\mathrm{BaCl_2}$ or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M $\mathrm{Na_2SO_4}$ or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume $\mathrm{Na_2SO_4}$ solution.
Set up and balance the chemical equation.	$\mathrm{Na_2SO_4(aq) + BaCl_2(aq) \longrightarrow BaSO_4(s) + \underline{2}NaCl(aq)}$ An insoluble product is formed after the reaction.
List other known quantities.	1 mol of $\mathrm{Na_2SO_4}$ to 1 mol $\mathrm{BaCl_2}$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$275\ \cancel{mL\ BaCl_2\ solution} \times \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \times \frac{0.250\ \cancel{mol\ BaCl_2}}{1\ \cancel{L\ BaCl_2\ solution}} \times \frac{1\ \cancel{mol\ Na}}{1\ \cancel{mol\ Ba}}$ = 0.1375 L sodium sulfate

Steps for Problem Solving	Example 1.9.7.1
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than barium chloride. Also, the units are correct.

? Exercise 1.9.7.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

Answer

0.250 L LiOH solution

1.9.7.1: Concentrations as Conversion Factors

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108\ L\ \cancel{NaCl} \times \frac{0.887\, mol\, NaCl}{L\ \cancel{NaCl}} = 0.0958\, mol\, NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 1.9.7.1.1

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88\ mol\ \cancel{CuSO_4} \times \frac{1\ L}{2.35\ \cancel{mol}} = 2.08\ L\ of\ solution$$

Exercise 1.9.7.1.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer:

1.74 L

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 1.9.7.1.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765\ \cancel{L} \times \frac{1.93\ mol\ \cancel{NaOH}}{L\ \cancel{solution}} \times \frac{40.0 g\, NaOH}{1\ mol\ \cancel{NaOH}} = 59.1\, g\, NaOH$$

? Exercise 1.9.7.1.2

What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer

6.12 g

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,AgNO_3(aq) + CaCl_2(aq) \rightarrow 2\,AgCl(s) + Ca(NO_3)_2(aq)$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25\ mol\ \cancel{AgNO_3} \times \frac{1\ mol\ \cancel{CaCl_2}}{2\ mol\ \cancel{AgNO_3}} \times \frac{1\,L\ solution}{0.555\ mol\ \cancel{CaCl_2}} = 1.13\ L\ CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

✓ Example 1.9.7.1.3

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,Ag(s) + Al(NO_3)_3(aq) \rightarrow 3\,AgNO_3 + Al(s)$$

Solution

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66\ g\ \cancel{Ag} \times \frac{1\ mol\ \cancel{Ag}}{107.97\ g\ \cancel{Ag}} \times \frac{1\ mol\ \cancel{Al(NO_3)_3}}{3\ mol\ \cancel{Ag}} \times \frac{1\,L\ solution}{0.0995\ mol\ \cancel{Al(NO_3)_3}} = 0.114\ L$$

The strikeouts show how the units cancel.

? Exercise 1.9.7.1.3

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4(s) + 2\,NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\,H_2O(\ell)$$

Answer:

0.777 L

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

✓ Example 1.9.7.1.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of $FeCl_3$. The student carefully adds 0.1074 M $Na_2C_2O_4$ until all the Fe^{3+}(aq) has precipitated as $Fe_2(C_2O_4)_3$(s). Using a precisely measured tube called a burette, the student finds that 9.04 mL of the $Na_2C_2O_4$ solution was added to completely precipitate the Fe^{3+}(aq). What was the concentration of the $FeCl_3$ in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,\mathrm{FeCl_3(aq)} + 3\,\mathrm{Na_2C_2O_4(aq)} \rightarrow \mathrm{Fe_2(C_2O_4)_3(s)} + 6\,\mathrm{NaCl(aq)}$$

Solution

First we need to determine the number of moles of $Na_2C_2O_4$ that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\ m\cancel{L} \times \frac{1\ \cancel{L}}{1000\ m\cancel{L}} \times \frac{0.1074 mol\ Na_2C_2O_4}{\cancel{L}} = 0.000971\ mol\ Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of Fe^{3+}(aq) that were present in the initial aliquot:

$$0.000971\ mol\ \cancel{Na_2C_2O_4} \times \frac{2mol\ FeCl_3}{3\ mol\cancel{Na_2C_2O_4}} = 0.000647 mol\ FeCl_3$$

Then we determine the concentration of $FeCl_3$ in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\ FeCl_3}{0.01000 L} = 0.0647 M\ FeCl_3$$

? Exercise 1.9.7.1.4

A student titrates 25.00 mL of H_3PO_4 with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of H_3PO_4?

$$\mathrm{H_3PO_4(aq)} + 3\,\mathrm{KOH(aq)} \rightarrow \mathrm{K_3PO_4(aq)} + 3\,\mathrm{H_2O}$$

Answer:

0.0711 M

Figure 1.9.7.1.1 Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

Example 1.9.7.1.5

H_2O_2 is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,\mathrm{MnO_4^-(aq)} + 5\,\mathrm{H_2O_2(aq)} + 6\,\mathrm{H^+(aq)} \rightarrow 2\,\mathrm{Mn_2^+(aq)} + 5\,\mathrm{O_2(g)} + 8\,\mathrm{H_2O}(\ell)$$

What mass of 3.00% m/m H_2O_2 solution is needed to react with 0.355 mol of MnO_4^-(aq)?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2. Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355\ mol\ \cancel{MnO_4^-} \times \frac{5\ mol\ \cancel{H_2O_2}}{2\ mol\cancel{MnO_4^-}} \times \frac{34.02\ g\,\cancel{H_2O_2}}{mol\ \cancel{H_2O_2}} \times \frac{100g\ solution}{3\ g\,\cancel{H_2O_2}} = 1006g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2, and the third conversion factor comes from the definition of percentage concentration by mass.

Exercise 1.9.7.1.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^-.

Answer

7.28 g

Summary

Know how to apply concentration units as conversion factors.

1.9.7.2: Practice Solution Concentration

Exercise 1.9.7.2.1

How many grams of glucose (molar mass = 180.18 g/mol) are dissolved in 175 mL of a 2.00 M solution?

Answer

63.1 g. (0.350 moles)

Exercise 1.9.7.2.1

How many mL of 5.25 M HCl solution would be needed to react with 6.00 g of Zn in the reaction below?

$Zn_{(s)} + 2\ HCl_{(aq)}$ --> $ZnCl_{2\,(aq)} + H_{2\,(g)}$

Answer

35.0 mL (0.0350 L, 0.184 moles)

Titration Problems

Exercise 1.9.7.2.1

You perform the titration below on 12.00 mL of H_2SO_4 solution. You find that it takes 26.44 mL of 0.4483 M NaOH to react completely. What is the concentration of the H_2SO_4 solution?

$H_2SO_{4\,(aq)} + 2\ NaOH_{(aq)} \rightarrow Na_2SO_{4\,(aq)} + 2\ H_2O_{(l)}$

Answer

0.4939 M H_2SO_4

You perform the titration below on 20.00 mL of HCl solution. You find that it takes 18.72 mL of 0.3786 M NaOH to react completely. What is the concentration of the HCl solution?

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

Answer

0.3544 M HCl

You perform the titration below on 15.00 mL of HCl solution. You find that it takes 26.44 mL of 0.1352 M $Ba(OH)_2$ to react completely. What is the concentration of the HCl solution?

$2\ HCl_{(aq)} + Ba(OH)_{2\,(aq)} \rightarrow BaCl_{2\,(aq)} + 2\ H_2O_{(l)}$

Answer

0.4766 M HCl

Solution Stoichiometry

Exercise 1.9.7.2.1

In the reaction below, what volume of 0.7644 M $Pb(NO_3)_2$ solution is needed in order to make 37.06 g $PbCl_2$? (Molar mass of $PbCl_2$ is 278.1 g/mol. NaCl is in excess.)

$Pb(NO_3)_{2\,(aq)} + 2\ NaCl_{(aq)} \rightarrow 2\ NaNO_{3\,(aq)} + PbCl_{2\,(s)}$

Answer

174.3 mL (0.1743 L)

In the reaction below, what volume of 0.3805 M $Pb(NO_3)_2$ solution is needed to react with 455.0 mL of 0.6322 M NaCl solution?

$Pb(NO_3)_{2\ (aq)} + 2\ NaCl_{\ (aq)} \rightarrow 2\ NaNO_{3\ (aq)} + PbCl_{2\ (s)}$

Answer

378.0 mL (0.3780 L)

1.9.8: Freezing Point Depression and Boiling Point Elevation

Note from Dr. B.

We are not going to be using the concentration molality (m, moles solute/kilogram solvent). However, the concentration of particles can be calculated with the i value and molarity as well. We will only be concerned with comparing two solutions' boiling or freezing points, not calculating the boiling or freezing point using the equation with molality.

Learning Objectives

- Explain what the term "colligative" means, and list the colligative properties.
- Indicate what happens to the boiling point and the freezing point of a solvent when a solute is added to it.
- Calculate boiling point elevations and freezing point depressions for a solution.

People who live in colder climates have seen trucks put salt on the roads when snow or ice is forecast. Why is this done? As a result of the information you explore in this section, you will understand why these events occur. You will also learn to calculate exactly how much of an effect a specific solute can have on the boiling point or freezing point of a solution.

The example given in the introduction is an example of a colligative property. **Colligative properties** are properties that differ based on the concentration of solute in a solvent, but not on the type of solute. What this means for the example above is that people in colder climates do not necessarily need salt to get the same effect on the roads—any solute will work. However, the higher the concentration of solute, the more these properties will change.

Boiling Point Elevation

Water boils at $100^{o}\mathrm{C}$ at $1\ \mathrm{atm}$ of pressure, but a solution of saltwater does not . When table salt is added to water, the resulting solution has a higher boiling point than the water did by itself. The ions form an attraction with the solvent particles that prevents the water molecules from going into the gas phase. Therefore, the saltwater solution will not boil at $100^{o}\mathrm{C}$. In order for the saltwater solution to boil, the temperature must be raised about $100^{o}\mathrm{C}$. This is true for any solute added to a solvent; the boiling point will be higher than the boiling point of the pure solvent (without the solute). In other words, when anything is dissolved in water, the solution will boil at a higher temperature than pure water would.

The boiling point elevation due to the presence of a solute is also a colligative property. That is, the amount of change in the boiling point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20\ \mathrm{m}$ solution of table salt and a $0.20\ \mathrm{m}$ solution of hydrochloric acid would have the same effect on the boiling point.

Freezing Point Depression

The effect of adding a solute to a solvent has the opposite effect on the freezing point of a solution as it does on the boiling point. A solution will have a lower freezing point than a pure solvent. The **freezing point** is the temperature at which the liquid changes to a solid. At a given temperature, if a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from going into the solid phase. The solute-solvent interactions require the temperature to decrease further in order to solidify the solution. A common example is found when salt is used on icy roadways. Salt is put on roads so that the water on the roads will not freeze at the normal $0^{o}\mathrm{C}$ but at a lower temperature, as low as $-9^{o}\mathrm{C}$. The de-icing of planes is another common example of freezing point depression in action. A number of solutions are used, but commonly a solution such as ethylene glycol, or a less toxic monopropylene glycol, is used to de-ice an aircraft. The aircrafts are sprayed with the solution when the temperature is predicted to drop below the freezing point. The **freezing point depression** is the difference in the freezing points of the solution from the pure solvent. This is true for any solute added to a solvent; the freezing point of the solution will be lower than the freezing point of the pure solvent (without the solute). Thus, when anything is dissolved in water, the solution will freeze at a lower temperature than pure water would.

The freezing point depression due to the presence of a solute is also a colligative property. That is, the amount of change in the freezing point is related to the number of particles of solute in a solution and is not related to the chemical composition of the solute. A $0.20\ \mathrm{m}$ solution of table salt and a $0.20\ \mathrm{m}$ solution of hydrochloric acid would have the same effect on the freezing point.

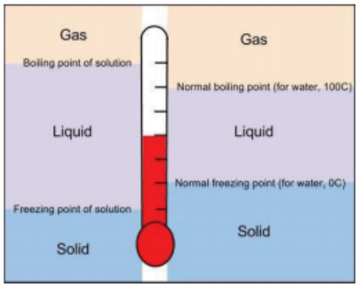


Figure 1.9.8.1: Comparison of boiling and freezing points of a pure liquid (right side) with a solution (left side).

Comparing the Freezing and Boiling Point of Solutions

Recall that covalent and ionic compounds do not dissolve in the same way. Ionic compounds break up into cations and anions when they dissolve. Covalent compounds typically do not break up. For example a sugar/water solution stays as sugar and water, with the sugar molecules staying as molecules. Remember that colligative properties are due to the number of solute particles in the solution. Adding 10 molecules of sugar to a solvent will produce 10 solute particles in the solution. When the solute is ionic, such as NaCl however, adding 10 formulas of solute to the solution will produce 20 ions (solute particles) in the solution. Therefore, adding enough NaCl solute to a solvent to produce a $0.20\ \mathrm{m}$ solution will have twice the effect of adding enough sugar to a solvent to produce a $0.20\ \mathrm{m}$ solution. Colligative properties depend on the number of solute particles in the solution.

"i" is the number of particles that the solute will dissociate into upon mixing with the solvent. For example, sodium chloride, NaCl, will dissociate into two ions so for NaCl, $i=2$; for lithium nitrate, $\mathrm{LiNO_3}$, $i=2$; and for calcium chloride, $\mathrm{CaCl_2}$, $i=3$. For covalent compounds, i is always equal to 1.

By knowing the *molality* of a solution and the number of particles a compound will dissolve to form, it is possible to predict which solution in a group will have the lowest freezing point. To compare the boiling or freezing points of solutions, follow these general steps:

1. Label each solute as ionic or covalent.
2. If the solute is ionic, determine the number of *ions* in the formula. Be careful to look for polyatomic ions.
3. Multiply the original molality (m) of the solution by the number of particles formed when the solution dissolves. This will give you the total concentration of particles dissolved.
4. Compare these values. The higher total concentration will result in a higher boiling point and a lower freezing point.

✓ Example 1.9.8.1

Rank the following solutions in water in order of *increasing* (lowest to highest) freezing point:

- $0.1\ \mathrm{m\ NaCl}$
- $0.1\ \mathrm{m\ C_6H_{12}O_6}$
- $0.1\ \mathrm{m\ CaI_2}$

Solution

To compare freezing points, we need to know the total concentration of all particles when the solute has been dissolved.

- $0.1\ \mathrm{m\ NaCl}$: This compound is ionic (metal with nonmetal), and will dissolve into 2 parts. The total final concentration is: $(0.1\ \mathrm{m})(2)=0.2\ \mathrm{m}$
- $0.1\ \mathrm{m\ C_6H_{12}O_6}$: This compound is covalent (nonmetal with nonmetal), and will stay as 1 part. The total final concentration is: $(0.1\ \mathrm{m})(1)=0.1\ \mathrm{m}$
- $0.1\ \mathrm{m\ CaI_2}$: This compound is ionic (metal with nonmetal), and will dissolve into 3 parts. The total final concentration is: $(0.1\ \mathrm{m})(3)=0.3\ \mathrm{m}$

Remember, the greater the concentration of particles, the lower the freezing point will be. $0.1\ \mathrm{m\ CaI_2}$ will have the lowest freezing point, followed by $0.1\ \mathrm{m\ NaCl}$, and the highest of the three solutions will be $0.1\ \mathrm{m\ C_6H_{12}O_6}$, but all three of them will have a lower freezing point than pure water.

The boiling point of a solution is higher than the boiling point of a pure solvent, and the freezing point of a solution is lower than the freezing point of a pure solvent. However, the amount to which the boiling point increases or the freezing point decreases depends on the amount of solute that is added to the solvent. A mathematical equation is used to calculate the boiling point elevation or the freezing point depression.

The boiling point elevation is the amount that the boiling point temperature *increases* compared to the original solvent. For example, the boiling point of pure water at 1.0 atm is 100^oC while the boiling point of a 2% saltwater solution is about 102^oC. Therefore, the boiling point elevation would be 2^oC. The freezing point depression is the amount that the freezing temperature *decreases*.

Both the boiling point elevation and the freezing point depression are related to the molality of the solution. Looking at the formula for the boiling point elevation and freezing point depression, we see similarities between the two. The equation used to calculate the increase in the boiling point is:

$$\Delta T_b = k_b \cdot \text{m} \cdot i \tag{1.9.8.1}$$

Where:

- $\Delta T_b =$ the amount the boiling point increases.
- $k_b =$ the boiling point elevation constant which depends on the solvent (for water, this number is $0.515^\text{o}\text{C}/\text{m}$).
- $\text{m} =$ the molality of the solution.
- $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1).

The following equation is used to calculate the decrease in the freezing point:

$$\Delta T_f = k_f \cdot \text{m} \cdot i \tag{1.9.8.2}$$

Where:

- $\Delta T_f =$ the amount the freezing temperature decreases.
- $k_f =$ the freezing point depression constant which depends on the solvent (for water, this number is $1.86^\text{o}\text{C}/\text{m}$).
- $\text{m} =$ the molality of the solution.
- $i =$ the number of particles formed when that compound dissolves (for covalent compounds, this number is always 1).

✓ Example 1.9.8.2: Adding Antifreeze to Protein Engines

Antifreeze is used in automobile radiators to keep the coolant from freezing. In geographical areas where winter temperatures go below the freezing point of water, using pure water as the coolant could allow the water to freeze. Since water expands when it freezes, freezing coolant could crack engine blocks, radiators, and coolant lines. The main component in antifreeze is ethylene glycol, $\text{C}_2\text{H}_4(\text{OH})_2$. What is the concentration of ethylene glycol in a solution of water, in molality, if the freezing point dropped by 2.64^oC? The freezing point constant, k_f, for water is $1.86^\text{o}\text{C}/\text{m}$.

Solution

Use the equation for freezing point depression of solution (Equation 1.9.8.2):

$$\Delta T_f = k_f \cdot \text{m} \cdot i$$

Substituting in the appropriate values we get:

$$2.64^\text{o}\text{C} = (1.86^\text{o}\text{C}/\text{m})\,(\text{m})\,(1)$$

Solve for m by dividing both sides by $1.86^\text{o}\text{C}/\text{m}$.

$$\text{m} = 1.42$$

✓ Example 1.9.8.3: Adding Salt to Elevate Boiling Temperature

A solution of 10.0 g of sodium chloride is added to 100.0 g of water in an attempt to elevate the boiling point. What is the boiling point of the solution? k_b for water is $0.52^\text{o}\text{C}/\text{m}$.

Solution

Use the equation for boiling point elevation of solution (Equation 1.9.8.1):

$$\Delta T_b = k_b \cdot \text{m} \cdot i$$

We need to be able to substitute each variable into this equation.

- $k_b = 0.52^{\text{o}}\text{C}/\text{m}$
- m: We must solve for this using stoichiometry. Given: $10.0\ \text{g NaCl}$and $100.0\ \text{g H}_2\text{O}$Find: $\text{mol NaCl}/\text{kg H}_2\text{O}$. Ratios: molar mass of NaCl, $1000\ \text{g} = 1\ \text{kg}$

$$\frac{10.0\ \ \text{g}\ \cancel{\text{NaCl}}}{100.0\ \ \text{g}\ \cancel{\text{H}_2\text{O}}} \cdot \frac{1\ \text{mol NaCl}}{58.45\ \ \text{g}\ \cancel{\text{NaCl}}} \cdot \frac{1000\ \ \text{g}\ \cancel{\text{H}_2\text{O}}}{1\ \text{kg H}_2\text{O}} = 1.71\ \text{m}$$

- For NaCl, $i = 2$

Substitute these values into the equation $\Delta T_b = k_b \cdot \text{m} \cdot i$. We get:

$$\Delta T_b = \left(0.52 \frac{^{\text{o}}\text{C}}{\cancel{\text{m}}}\right)(1.71\ \ \cancel{\text{m}})\,(2) = 1.78^{\text{o}}\text{C}$$

Water normally boils at 100^{o}C, but our calculation shows that the boiling point increased by 1.78^{o}C. Our new boiling point is $101.78^{\text{o}}\text{C}$.

Note: Since sea water contains roughly 28.0 g of NaCl per liter, this saltwater solution is approximately **four times** more concentrated than sea water (all for a 2° C rise of boiling temperature).

Summary

- Colligative properties are properties that are due only to the number of particles in solution, and are not related to the chemical properties of the solute.
- Boiling points of solutions are higher than the boiling points of the pure solvents.
- Freezing points of solutions are lower than the freezing points of the pure solvents.
- Ionic compounds split into ions when they dissolve, forming more particles. Covalent compounds stay as complete molecules when they dissolve.

Vocabulary

- **Colligative property** - A property that is due only to the number of particles in solution, and not the type of the solute.
- **Boiling point elevation** - The amount that the boiling point of a solution increases from the boiling point of the pure solvent.
- **Freezing point depression** - The amount that the freezing point of a solution decreases from the freezing point of the pure solvent.

1.9.8.1: Osmosis

Learning Objectives

- Explain the following laws within the Ideal Gas Law

Before we introduce the final colligative property, we need to present a new concept. A **semipermeable membrane** is a thin membrane that will pass certain small molecules, but not others. A thin sheet of cellophane, for example, acts as a semipermeable membrane. Consider the system in Figure 1.9.8.1.1

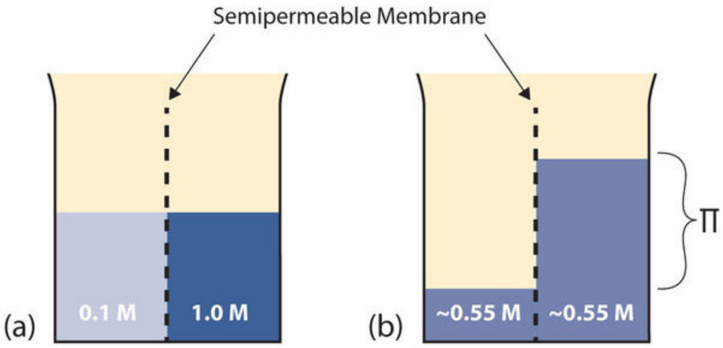


Figure 1.9.8.1.1: Osmosis. (a) Two solutions of differing concentrations are placed on either side of a semipermeable membrane. (b) When osmosis occurs, solvent molecules selectively pass through the membrane from the dilute solution to the concentrated solution, diluting it until the two concentrations are the same. The pressure exerted by the different height of the solution on the right is called the osmotic pressure. (CC BY-SA-NC 3.0; anonymous)

a. A semipermeable membrane separates two solutions having the different concentrations marked. Curiously, this situation is not stable; there is a tendency for water molecules to move from the dilute side (on the left) to the concentrated side (on the right) until the concentrations are equalized, as in Figure 1.9.8.1.1b
b. This tendency is called **osmosis**. In osmosis, the solute remains in its original side of the system; only solvent molecules move through the semipermeable membrane. In the end, the two sides of the system will have different volumes. Because a column of liquid exerts a pressure, there is a pressure difference (Π) on the two sides of the system that is proportional to the height of the taller column. This pressure difference is called the **osmotic pressure**, which is a colligative property.

The osmotic pressure of a solution is easy to calculate:

\[\Pi = MRT \nonumber \]

where Π is the osmotic pressure of a solution, *M* is the molarity of the solution, *R* is the ideal gas law constant, and *T* is the absolute temperature. This equation is reminiscent of the ideal gas law we considered in Chapter 6.

Example 1.9.8.1.5: Osmotic Pressure

What is the osmotic pressure of a 0.333 M solution of $C_6H_{12}O_6$ at 25°C?

Solution

First we need to convert our temperature to kelvins:

$$T = 25 + 273 = 298\ \text{K}$$

Now we can substitute into the equation for osmotic pressure, recalling the value for *R*:

$$\Pi = (0.333 M)\left(0.08205 \frac{L.atm}{mol.K}\right)(298 K)$$

The units may not make sense until we realize that molarity is defined as moles per liter:

$$\Pi = \left(0.333 \frac{mol}{L}\right)\left(0.08205 \frac{L.atm}{mol.K}\right)(298 K)$$

Now we see that the moles, liters, and kelvins cancel, leaving atmospheres, which is a unit of pressure. Solving,

$$\Pi = 8.14\, atm$$

This is a substantial pressure! It is the equivalent of a column of water 84 m tall.

Exercise 1.9.8.1.5

What is the osmotic pressure of a 0.0522 M solution of $C_{12}H_{22}O_{11}$ at 55°C?

Answer

1.40 atm

Osmotic pressure is important in biological systems because cell walls are semipermeable membranes. In particular, when a person is receiving intravenous (IV) fluids, the osmotic pressure of the fluid needs to be approximately the same as blood serum to avoid any negative consequences. Figure 1.9.8.1.3shows three red blood cells:

- A healthy red blood cell.
- A red blood cell that has been exposed to a lower concentration than normal blood serum (a *hypotonic* solution); the cell has plumped up as solvent moves into the cell to dilute the solutes inside.
- A red blood cell exposed to a higher concentration than normal blood serum (*hypertonic*); water leaves the red blood cell, so it collapses onto itself. Only when the solutions inside and outside the cell are the same (*isotonic*) will the red blood cell be able to do its job.

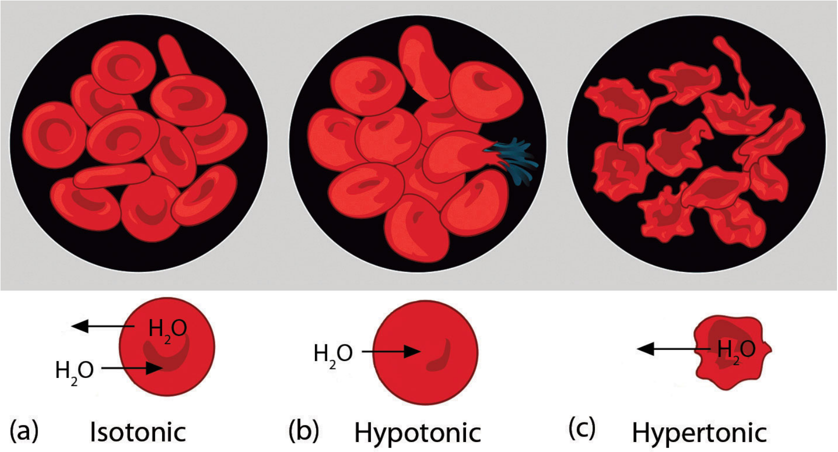


Figure 1.9.8.1.3 : Osmotic Pressure and Red Blood Cells. (a) This is what a normal red blood cell looks like. (b) When a red blood cell is exposed to a hypotonic solution, solvent goes through the cell membrane and dilutes the inside of the cell. (c) When a red blood cell is exposed to a hypertonic solution, solvent goes from the cell to the surrounding solution, diluting the hypertonic solution and collapsing the cell. Neither of these last two cases is desirable, so IV solutions must be isotonic with blood serum to not cause deleterious effects. (Public Domain; Mariana Ruiz Villareal)

Osmotic pressure is also the reason you should not drink seawater if you're stranded on a lifeboat in the ocean; seawater has a higher osmotic pressure than most of the fluids in your body. You *can* drink the water, but ingesting it will pull water out of your cells as osmosis works to dilute the seawater. Ironically, your cells will die of thirst, and you will also die. (It is okay to drink the water if you are stranded on a body of freshwater, at least from an osmotic pressure perspective.) Osmotic pressure is also thought to be important—in addition to capillary action—in getting water to the tops of tall trees.

Summary

- Osmotic pressure is caused by concentration differences between solutions separated by a semipermeable membrane, and is an important biological consideration.

1.9.8.2: Practice Colligative Properties

Exercise 1.9.8.2.1

A red blood cell is isotonic with a 0.9 % (m/v) NaCl solution. If you placed red blood cells into 100.0 mL of solution that has 2.0 g of NaCl dissolved in it, what will happen to the cells?

Answer

The NaCl solution is 2.0 % (m/v), so it is hypertonic compared to the cell. (The cell is hypotonic compared to the solution.) When the cell is placed into the solution, water will flow from the low concentration cell into the high concentration salt water. The cell will shrivel (crenation).

Exercise 1.9.8.2.1

Which solution will have a lower freezing point, one with 25.0 g of NaCl dissolved in 0.500 L solution, or one with 25.0 g of $MgCl_2$ dissolved in 0.500 L solution? (Hint: what is molarity of each solution? What is number of particles "i"? What is molarity of particles in the solution?)

Answer

NaCl solution is 0.856 M, i=2, it is 1.71 M in particles.

$MgCl_2$ solution is 0.263 M, i=3, it is 0.789 M in particles.

The more concentrated NaCl will experience more freezing point depression and have a lower freezing point.

SECTION OVERVIEW

1.10: Equilibrium

1.10.1: The Rate of a Chemical Reaction

Learning Objectives

- Describe the conditions for successful collisions that cause reactions.
- Describe *rate* in terms of the conditions of successful collisions.
- Describe how changing the temperature, concentration of a reactant, or surface area of a reaction affects the rate of a reaction.
- Define a catalyst and how a catalyst affects the rate of a reaction.

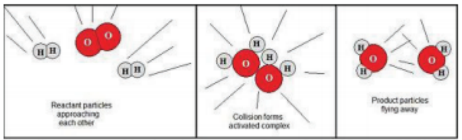
We know that a chemical system can be made up of atoms (H_2, N_2, K, etc.), ions (NO_3^-, Cl^-, Na^+, etc.), or molecules (H_2O, $C_{12}H_{22}O_{11}$, etc.). We also know that in a chemical system, these particles are moving around in random motion. The **collision theory** explains why reactions occur at this particle level between these atoms, ions, and/or molecules. It also explains how it is possible to speed up or slow down reactions that are occurring.

Collision Theory

The collision theory provides us with the ability to predict what conditions are necessary for a successful reaction to take place. These conditions include:

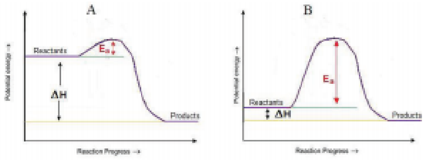
1. The particles must collide with each other.
2. The particles must collide with sufficient energy to break the old bonds.
3. The particles must have proper orientation.

A chemical reaction involves breaking bonds in the reactants, rearranging the atoms into new groupings (the products), and forming new bonds in the products.



This collision is successful and results in reaction. (Source: Richard Parsons. CC-BY-SA)

Therefore, a collision must not only occur between reactant particles, but the collision also has to have sufficient energy to break all the reactant bonds that need to be broken in order to form the products. Some reactions need less collision energy than others. The amount of energy that reactant particles must have in order to break the old bonds for a reaction to occur is called the **activation energy**, abbreviated E_a. Another way to think of this is to look at an energy diagram, as shown in the figure. Particles must be able to get over the "bump"—the activation energy—if they are going to react. If the reactant particles collide with less than the activation energy, the particles will rebound (bounce off of each other), and no reaction will occur.



Each reaction has its own activation energy, E_a. The smaller the "bump", the less energy particles must have to react.

Reaction Rate

Chemists use reactions to generate a product for which they have a use. For the most part, the reactions that produce some desired compound are only useful if the reaction occurs at a reasonable rate. For example, using a reaction to produce brake fluid would not be useful if the reaction required 8,000 years to complete the product. Such a reaction would also not be useful if the reaction was so fast that it was explosive. For these reasons, chemists wish to be able to control reaction rates. In some cases, chemists wish to speed up reactions that are too slow or slow down reactions that are too fast. In order to gain any control over reaction rates, we must know the factors that affect reaction rates. Chemists have identified many factors that affect the rate of a reaction.

The rate, or speed, at which a reaction occurs depends on the frequency of successful collisions. Remember, a successful collision occurs when two reactants collide with enough energy and with the right orientation. That means if there is an increase in the number of collisions, an increase in the number of particles that have enough energy to react, and/or an increase in the number of particles with the correct orientation, the rate of reaction will increase.

Effect of Temperature on Rate of Reaction

The rate of reaction was discussed in terms of three factors: collision frequency, the collision energy, and the geometric orientation. Remember that the collision frequency is the number of collisions per second. The collision frequency is dependent, among other factors, on the temperature of the reaction.

When the temperature is increased, the average velocity of the particles is increased. The average kinetic energy of these particles is also increased. The result is that the particles will collide more frequently, because the particles move around faster and will encounter more reactant particles. However, this is only a minor part of the reason why the rate is increased. Just because the particles are colliding more frequently does not mean that the reaction will definitely occur.

The major effect of increasing the temperature is that more of the particles that collide will have the amount of energy needed to have an effective collision. In other words, more particles will have the necessary activation energy.

At room temperature, the hydrogen and oxygen in the atmosphere do not have sufficient energy to attain the activation energy needed to produce water:

$$\mathrm{O_2}\,(g)+\mathrm{H_2}\,(g)\rightarrow \text{No reaction}$$

At any one moment in the atmosphere, there are many collisions occurring between these two reactants. But what we find is that water is not formed from the oxygen and hydrogen molecules colliding in the atmosphere, because the activation energy barrier is just too high, and all the collisions are resulting in rebound. When we increase the temperature of the reactants or give them energy in some other way, the molecules have the necessary activation energy and are able to react to produce water:

$$\mathrm{O_2}\,(g)+\mathrm{H_2}\,(g)\rightarrow \mathrm{H_2O}\,(l)$$

There are times when the rate of a reaction needs to be slowed down. Lowering the temperature could also be used to decrease the number of collisions that would occur and lowering the temperature would also reduce the kinetic energy available for activation energy. If the particles have insufficient activation energy, the collisions will result in rebound rather than reaction. Using this idea, when the rate of a reaction needs to be lower, keeping the particles from having sufficient activation energy will definitely keep the reaction at a lower rate.

Society uses the effects of temperature on reaction rate every day. Food storage is a prime example of how the temperature effect on reaction rate is utilized by society. Consumers store food in freezers and refrigerators to slow down the processes that cause it to spoil. The decrease in temperature decreases the rate at which food will break down or be broken down by bacteria. In the early years of the 20^{th} century, explorers were fascinated with being the first to reach the South Pole. In order to attempt such a difficult task at a time without most of the technology that we take for granted today, they devised a variety of ways of surviving. One method was to store their food in the snow to be used later during their advances to the pole. On some explorations, they buried so much food that they didn't need to use all of it, and some was left behind. Many years later, when this food was located and thawed, it was found to still be edible.

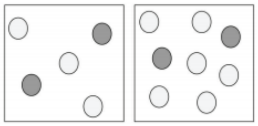
When milk, for example, is stored in the refrigerator, the molecules in the milk have less energy. This means that while molecules will still collide with other molecules, few of them will react (which means in this case "spoil") because the molecules do not have sufficient energy to overcome the activation energy barrier. The molecules do have energy and are colliding, however, and so, over time, even in the refrigerator, the milk will spoil. Eventually the higher energy molecules will gain the energy needed to react and when enough of these reactions occur, the milk becomes "soured".

However, if that same carton of milk was at room temperature, the milk would react (in other words, "spoil") much more quickly. Most of the molecules would have sufficient energy to overcome the energy barrier at room temperature, and many more collisions would occur. This allows for the milk to spoil in a fairly short amount of time. This is also the reason why most fruits and vegetables ripen in the summer when the temperature is much warmer. You may have experienced this first hand if you have ever bitten into an unripe banana—it was probably sour tasting and might even have felt like biting into a piece of wood! When a banana ripens, numerous reactions occur that produce all the compounds that we expect to taste in a banana. But this can only happen if the temperature is high enough to allow these reactions to make those products.

Effect of Concentration on Rate of Reaction

If you had an enclosed space, like a classroom, and there was one red ball and one green ball flying around the room in random motion, undergoing perfectly elastic collisions with the walls and with each other, in a given amount of time, the balls would collide with each other a certain number of times determined by probability. If you now put two red balls and one green ball in the room under the same conditions, the probability of a collision between a red ball and the green ball would exactly double. The green ball would have twice the chance of encountering a red ball in the same amount of time.

In terms of chemical reactions, a similar situation exists. Particles of two gaseous reactants or two reactants in solution have a certain probability of undergoing collisions with each other in a reaction vessel. If you double the concentration of either reactant, the probability of a collision doubles. The rate of reaction is proportional to the number of collisions per unit time. If one concentration is doubled, the number of collisions will also double. Assuming that the percent of collisions that are successful does not change, then having twice as many collisions



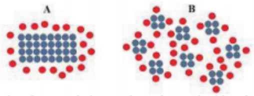
The reaction mixture on the left is less concentrated, so particles will not collide as often. The reaction will be slower.
Because the reacting particles on the right image have a greater change of colliding, the reaction will go faster.
CC – Tracy Poulsen

will result in twice as many successful collisions. The rate of reaction is proportional to the number of collisions over time; increasing the concentration of either reactant increases the number of collisions, and therefore increases the number of successful collisions and the reaction rate.

For example, the chemical test used to identify a gas as oxygen, or not, relies on the fact that increasing the concentration of a reactant increases reaction rate. The reaction we call combustion refers to a reaction in which a flammable substance reacts with oxygen. If we light a wooden splint (a thin splinter of wood) on fire and then blow the fire out, the splint will continue to glow in air for a period of time. If we insert that glowing splint into any gas that does not contain oxygen, the splint will immediately cease to glow—that is, the reaction stops. Oxygen is the only gas that will support combustion, Air is approximately 20% oxygen gas. If we take that glowing splint and insert it into pure oxygen gas, the reaction will increase its rate by a factor of five, since pure oxygen has 5 times the concentration of oxygen that is in the air. When the reaction occurring on the glowing splint increases its rate by a factor of five, the glowing splint will suddenly burst back into full flame. This test, of thrusting a glowing splint into a gas, is used to identify the gas as oxygen. Only a greater concentration of oxygen than that found in air will cause the glowing splint to burst into flame.

Effect of Surface Area on Rate of Reaction

The very first requirement for a reaction to occur between reactant particles is that the particles must collide with one another. The previous section pointed out how increasing the concentration of the reactants increases reaction rate because it increases the frequency of collisions between particles. It can be shown that the number of collisions that occur between reactant particles is also dependent on the surface area of solid reactants. Consider a reaction between reactant RED and reactant BLUE in which reactant blue is in the form of a single lump. Then compare this to the same reaction where reactant blue has been broken up into many smaller pieces.



In these figures, only the particles on the outside of the solid blue reactant have a chance to collide with the red reactant. In figure B, the same amount of solid reactant as used in A was crushed into smaller particles. This means that more particles on the outside of the reactant have an opportunity to collide with the red reactant and speeds up the reaction.

In the diagram, only the blue particles on the outside surface of the lump are available for collision with reactant red. The blue particles on the interior of the lump are protected by the blue particles on the surface. In Figure A, if you count the number of blue particles available for collision, you will find that only 20 blue particles could be struck by a particle of reactant red. In Figure A, there are a number of blue particles on the interior of the lump that cannot be struck. In Figure B, however, the lump has been broken up into smaller pieces and all the interior blue particles are now on a surface and available for collision. In Figure B, more collisions between the blue and red will occur, and therefore, the reaction in Figure B will occur at a faster rate than the same reaction in Figure A. Increasing the surface area of a reactant increases the frequency of collisions and increases the reaction rate.

Several smaller particles have more surface area than one large particle. The more surface area that is available for particles to collide, the faster the reaction will occur. You can see an example of this in everyday life if you have ever tried to start a fire in the fireplace. If you hold a match up against a large log in an attempt to start the log burning, you will find it to be an unsuccessful effort. Holding a match against a large log will not cause enough reactions to occur in order to keep the fire going by providing sufficient activation energy for further reactions. In order to start a wood fire, it is common to break a log up into many small, thin sticks called kindling. These thinner sticks of wood provide many times the surface area of a single log. The match will successfully cause enough reactions in the kindling so that sufficient heat is given off to provide activation energy for further reactions.

There have been, unfortunately, cases where serious accidents were caused by the failure to understand the relationship between surface area and reaction rate. One such example occurred in flour mills. A grain of wheat is not very flammable. It takes significant effort to get a grain of wheat to burn. If the grain of wheat, however, is pulverized and scattered through the air, only a spark is necessary to cause an explosion. When the wheat is ground to make flour, it is pulverized into a fine powder and some of the powder gets scattered around in the air. A small spark then, is sufficient to start a very rapid reaction which can destroy the entire flour mill. In a 10-year period from 1988 to 1998, there were 129 grain dust explosions in mills in the United States. Efforts are now made in flour mills to have huge fans circulate the air in the mill through filters to remove the majority of the flour dust particles.

Another example is in the operation of coal mines. Coal will of course burn, but it takes an effort to get the coal started; once it is burning, it burns slowly because only the surface particles are available to collide with oxygen particles. The interior particles of coal have to wait until the outer surface of the coal lump burns off before they can collide with oxygen. In coal mines, huge blocks of coal must be broken up before the coal can be brought out of the mine. In the process of breaking up the huge blocks of coal, drills are used to drill into the walls of coal. This drilling produces fine coal dust that mixes into the air; then a spark from a tool can cause a massive explosion in the mine. There are explosions in coal mines for other reasons, but coal dust explosions have contributed to the death of many miners. In modern coal mines, lawn sprinklers are used to spray water through the air in the mine and this reduces the coal dust in the air, and eliminates coal dust explosions.

Effect of a Catalyst on Rate of Reaction

The final factor that affects the rate of the reaction is the effect of a catalyst. A **catalyst** is a substance that speeds up the rate of the reaction, without being consumed by the reaction itself.

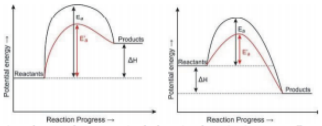
In the reaction of potassium chlorate breaking down to potassium chloride and oxygen, a catalyst is available to make this reaction occur much faster than it would occur by itself under room conditions. The reaction is:

$$2\text{KClO}_3\,(s)\overset{\text{MnO}_2(s)}{\longrightarrow}2\text{KCl}\,(s)+3\text{O}_2\,(g)$$

The catalyst is manganese dioxide and its presence causes the reaction shown above to run many times faster than it occurs without the catalyst. When the reaction has reached completion, the MnO_2 can be removed from the reaction vessel and its condition is exactly the same as it was before the reaction. This is part of the definition of a catalyst—that it is not consumed by the reaction. You should note that the catalyst is not written into the equation as a reactant or a product, but is noted above the yields arrow. This is standard notation for the use of a catalyst.

Some reactions occur very slowly without the presence of a catalyst. In other words, the activation energy for these reactions is very high. When the catalyst is added, the activation energy is lowered because the catalyst provides a new reaction pathway with lower activation energy.

In the figure on the right, the endothermic reaction shows the catalyst reaction in red with the lower activation energy, designated E'_a. The new reaction pathway has lower activation energy but has no effect on the energy of the reactants, the products, or the value of ΔH. The same is true for the exothermic reaction. The activation energy of the catalyzed reaction is lower than that of the uncatalyzed reaction. The new reaction pathway provided by the catalyst affects the energy required for reactant bonds to break, and product bonds to form.



A catalyst speeds up a reaction by lowering the activation energy, Ea,. Because less energy is required to react, more particles have the necessary energy.

While many reactions in the laboratory can be increased by increasing the temperature, this is not possible for all of the reactions that occur in our bodies throughout our entire lives. In fact, the body needs to be maintained at a very specific temperature: 98.6^oF or 37^oC. Of course there are times, such as when the body is fighting infection, when the body temperature may be increased. But generally, in a healthy person, the temperature is quite consistent. However, many of the reactions that a healthy body depends on could never occur at body temperature. The answer to this dilemma is catalysts—also referred to as enzymes. Many of these enzymes are made in human cells because human DNA carries the directions to make them. However, there are some enzymes required by the body that are not made by human cells. These catalysts must be supplied to our bodies in the food we eat and are called vitamins.

Reversible Reactions

Typically when we think of a chemical reaction, we think of the reactants getting totally used up so that none are left, and that we end up with only products. Also, we generally consider chemical reactions as one-way events. You may well have learned during earlier science classes that this is one way to distinguish chemical change from physical change—physical changes (such as the melting and freezing of ice) are easily reversed, but chemical changes cannot be reversed (pretty tough to un-fry an egg).

Throughout this chaper, we will see that this isn't always the case. We will see that many chemical reactions are, in fact, reversible under the right conditions. And because many reactions can be reversed, our idea of a reaction ending with no reactants left, only products, will need to be modified.

Here are some examples of reactions that can be reversed:

1.	Nitrogen dioxide, NO_2, a reddish-brown gas, reacts to form colorless dinitrogen tetroxide, N_2O_4 : $2\,\text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$ But the reaction can also go the other way—dinitrogen tetroxide also readily breaks down to form nitrogen dioxide: $\text{N}_2\text{O}_4(\text{g}) \rightarrow 2\,\text{NO}_2(\text{g})$ We typically write a reaction that can go in both directions by using a double arrow (which will sometimes appear as ↔ in these online notes): $2\,\text{NO}_2(\text{g}) \leftrightarrow \text{N}_2\text{O}_4(\text{g})$ Because the reaction continues in both directions at the same time, we never run out of either NO_2 or N_2O_4 . NO_2 is continually being used up to form N_2O_4, but at the same time N_2O_4 is forming more NO_2

2\. When hydrogen gas is passed over heated iron oxide, iron and steam are produced:

(1)	$Fe_3O_4(s) + 4\,H_2(g) \rightarrow 3\,Fe(s) + 4$
The reverse reaction can occur when steam is passed over red-hot iron:	
(2)	$3\,Fe(s) + 4\,H_2O(g) \rightarrow Fe_3O_4(s) +$
We can write these two equations together as:	
(3)	\(\ce{Fe_3O_4(s) + 4H_2(g) \leftrightarrow 3Fe(s) + 4H_2O(g)}\)

When we have a reversible reaction written in this way, we need to be able to distinguish between which way the reaction is headed. As written above in Reaction (3), we would say that in the ***forward reaction***, iron oxide and hydrogen gas, the reactants, produce the products iron and steam.
During the ***reverse reaction***, iron reacts with steam to produce the products iron oxide and hydrogen gas.
It is important to understand the terminology, and to use the terms correctly.
Does it matter which way we write our reversible reaction? It can also be written as

$$3\,Fe(s) + 4\,H_2O(g) \leftrightarrow Fe_3O_4(s) + 4\,H_2(g)$$

Now iron and steam are reactants of the forward direction, and iron oxide and hydrogen gas would be the reactants of the reverse direction.

Summary

- The collision theory explains why reactions occur between atoms, ions, and molecules.
- In order for a reaction to be effective, particles must collide with enough energy, and have the correct orientation.
- With an increase in temperature, there is an increase in energy that can be converted into activation energy in a collision, and that will increase the reaction rate. A decrease in temperature would have the opposite effect.
- With an increase in temperature, there is an increase in the number of collisions.
- Increasing the concentration of a reactant increases the frequency of collisions between reactants and will, therefore, increase the reaction rate.
- Increasing the surface area of a reactant (by breaking a solid reactant into smaller particles) increases the number of particles available for collision and will increase the number of collisions between reactants per unit time.
- A catalyst is a substance that speeds up the rate of the reaction without being consumed by the reaction itself. When a catalyst is added, the activation energy is lowered because the catalyst provides a new reaction pathway with lower activation energy.

Vocabulary

- **Catalyst** - A substance that speeds up the rate of the reaction without being consumed by the reaction itself.
- **Surface area to volume ratio** - The comparison of the volume inside a solid to the area exposed on the surface.

Further Reading/Supplemental Links

- Activation Energy: http://www.mhhe.com/physsci/chemistr...sh/activa2.swf
- learner.org/resources/series61.html The ***learner.org*** website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge. The website has one video that relates to this lesson called ***Molecules in Action***.
- www.vitamins-guide.net
- en.Wikipedia.org/wiki
- Observing molecules during chemical reactions helps explain the role of catalysts. Dynamic equilibrium is also demonstrated. Molecules in Action (www.learner.org/vod/vod_window.html?pid=806)
- Surface science examines how surfaces react with each other at the molecular level. On the Surface (www.learner.org/vod/vod_window.html?pid=812)

1.10.1.1: Practice Effects on Rate

Exercise 1.10.1.1.1

$C_2H_4\,_{(g)} + H_2\,_{(g)}$ --> $C_2H_6\,_{(g)}$

Which of the following will increase the rate of the reaction above? (Choose all that apply.)

a. adding H_2 gas

b. adding C_2H_6 gas

c. adding a catalyst

d. decreasing the temperature

Answer

a and c.

Exercise 1.10.1.1.1

According to collision theory, why would increasing the temperature of a reaction increase the reaction rate? (Choose all that apply.)

a. It would increase the space between molecules.

b. It would increase the energy of the collisions between molecules.

c. It would increase the frequency of collisions between molecules.

d. It would cause collisions to have a more favorable orientation.

Answer

b and c.

1.10.2: The Idea of Dynamic Chemical Equilibrium

Chapter 1 • Chapter 2 • Chapter 3 • Chapter 4 • Chapter 5 • Chapter 6 • Chapter 7 • Chapter 8 • Chapter 9
Chapter 10 • Chapter 11 • Chapter 12 • Chapter 13 • Chapter 14 • Chapter 15 • Chapter 16 • Chapter 17

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Learning Objectives

- Describe the three possibilities that exist when reactants come together.
- Describe what is occurring in a system at equilibrium.

Think for a minute about sitting down to a table to eat dinner. There are three possibilities that could happen when you eat dinner. You could (1) finish your entire dinner, (2) you could not want any of it and leave it all on your plate, or (3) you could eat some of it and leave some of it. Reactions have the same possibilities. Reactions do not always proceed all the way from start to finish. You may have reactions that (1) go to completion so that at the end, the reaction vessel contains all products and only products. Some reactions (2) may not start at all, so at the end the reaction vessel contains all reactants and only reactants. And some reactions (3) may start but not go to completion, that is, the reaction might start but not go completely to products. In this last case, the reaction vessel would contain some reactants and some products. In this section, we are going to take a closer look at the third type of reaction.

Reversible Reactions and Equilibrium

Consider the hypothetical reaction:

$$\mathrm{A + B \rightarrow C + D.}$$

If we looked at this reaction using what we have learned, this reaction will keep going, forming C and D until A and B run out. This is what we call an "irreversible reaction" or a "reaction that goes to completion".

Some reactions, however, are **reversible**, meaning the reaction can go backwards in which products react to form reactants, so that: $\mathrm{A + B \leftarrow C + D}$. The direction of the arrow shows that C and D are reacting to form A and B. What if the two reactions, the forward reaction and the reverse reaction, were occurring at the same time? What would this look like? If you could peer into the reaction, you would be able to find A, B, C, and D particles. A and B would react to form C and D at the same time that C and D are reacting to form A and B.

If the forward and reverse reactions are happening at the same rate, the reaction is said to be at **equilibrium** or **dynamic equilibrium**. At this point, the concentrations of A, B, C, and D are not changing (or, are constant) and we would see no difference in our reaction container, but reactions are still occurring in both directions. It is important to point out that having constant amounts of reactants and products does NOT mean that the concentration of the reactants is equal to the concentration of the products. It means they are not changing. These reactions appear to have stopped before one of the reactants has run out.

Chemists use a double-headed arrow, $\rightleftharpoons$, to show that a reaction is at equilibrium. We would write the example reaction as:

$$\mathrm{A + B \rightleftharpoons C + D.}$$

The arrow indicates that both directions of the reaction are happening.

Another way to think about reversible and irreversible reactions is to compare them to two types of games of tag. Reversible reactions are in many ways like a traditional game of tag: the "it" person can become "not it" and somebody who is "not it" is tagged and becomes "it". In this way, it is a reversible change. It is also like a reaction at equilibrium, because overall no change is occurring. There is always a constant number of "it" people and "not it" people in the game. Also, having constant numbers of "it" and "not it" people in our game does not mean that the number of "it" people (reactants) is equal to the number of "not it" people (products). Furthermore, this is similar to equilibrium in that this game never truly ends (unless everybody gets tired of playing). The game could go on forever. We could write this as the following reversible reaction:

$$\text{"It"} \rightleftharpoons \text{"Not it"}$$

Irreversible reactions (those that only go in one direction from reactants to products and cannot reach a state of equilibrium) are more like a game of sharks and minnows. In sharks and minnows, almost everybody starts out as a minnow. Once tagged, they become a shark. However, the difference here is that once you are a shark you are always a shark; there is no way to go back to becoming a minnow. The game continues until everybody has been tagged and becomes a shark. This is similar to irreversible reactions in that the reactants turn into products, but can't change back. Furthermore, the reaction will proceed until the reactants have been used up and there are not any more left. We could write the reaction as:

$$\text{Minnow} \rightarrow \text{Shark}$$

Here's another example of a reversible reaction—dissolving salt in a beaker of water, described by the following reaction:

$$\text{NaCl(s)} \leftrightarrow \text{NaCl(aq)}$$

If you keep adding more and more solid salt, eventually you'll reach the point where no more salt dissolves, and the excess sits at the bottom of the beaker. At this point we have a saturated solution. Has the dissolving reaction stopped? It would appear so, but that's not the case (wouldn't that be too easy?).

What happens in our saturated solution, which has reached the point of **equilibrium**, is that both the forward

$$\text{NaCl(s)} \rightarrow \text{NaCl(aq)}$$

and reverse

$$\text{NaCl(aq)} \rightarrow \text{NaCl(s)}$$

reactions are still going on, but at the same ***rate***. This in effect cancels out any observable, or measurable, changes in our system. At the same rate that solid NaCl produces aqueous NaCl (dissolved salt), the dissolved salt is re-crystallizing to form more solid NaCl.

> - **Equilibrium** is the state at which the rate of the forward reaction equals the rate of the reverse reaction.
> - At the point of equilibrium, no more measurable or observable changes in the system can be noted.

It is important for you to understand that equilibrium means the ***rates*** of the forward and reverse reactions are equal; it does ***not*** mean that there are equal amounts of reactants and products present at equilibrium.

For example, the following reaction was allowed to come to the point of equilibrium, and concentrations of all reaction participants were measured at that time:

$$\text{H}_2\text{(g)} + \text{I}_2\text{(g)} \leftrightarrow 2\,\text{HI(g)}$$

At equilibrium:	$[\text{H}_2] = 0.022\text{ M}$ $[\text{I}_2] = 0.022\text{ M}$ $[\text{HI}] = 0.156\text{ M}$

For this particular reversible reaction, there is more HI at equilibrium (0.156 M) than there is of H_2 and I_2 (both at 0.022M). We say that the product side of the reaction is favored.

> *Equilibrium does not mean equal amounts at equilibrium!*

Here is an example to help you understand how equilibrium works: imagine yourself on a escalator that is going down. You start at the top (reactants) and end up at the bottom (products). But when you are partway down, you start walking up the escalator as it continues going down. If you match your rate of walking up to the same rate that the escalator is going down, you make no progress and appear to be at a standstill. To an observer it would look as if you and the escalator had come to a stop, when actually both upward and downward movements continue.

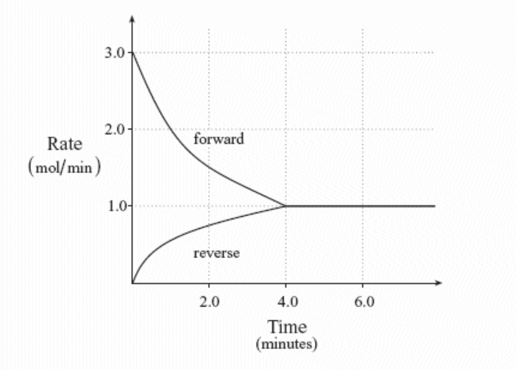
Equilibrium is **dynamic**—both forward and reverse reactions continue, even though the reaction appears to have stopped. And this equilibrium does not need to occur right in the middle of two floors—you could be near the bottom, near the top, or anywhere in between when you carry out your reverse process.

In order for a reversible reaction to reach the point of equilibrium, the reaction must be carried out in a closed system—no additional reactants can be added or products removed. If, in our last example, the product HI was removed as it formed, the reaction would never reach the point of equilibrium; instead, H_2 and I_2 would continue to react to produce HI until one or both of the reactants was used up.

If reactants are constantly being added, and products removed as they form, the system would appear to be at equilibrium because to an outside observer it would appear that the reaction has stopped—but that would not be the case. This situation—with new material constantly being added as products are removed—is called a **steady state system**. A factory with an assembly line is a steady state system—new raw materials are constantly being added; finished products are removed. A campfire with wood being added to the fire is another steady state system. Be careful not to confuse steady state with equilibrium.

How do the rates of the forward and reverse reactions change as the reaction heads towards equilibrium (before it reaches equilibrium)?

If we start our above reaction with H_2 and I_2, and with no HI, the two gases will react at a certain rate. But remember that the rate of a reaction slows down over time, as the reactants get used up (and lower their concentrations). Eventually, however, the amount of the product HI increases, and it will begin producing H_2 and I_2. Thus the rate of the reverse reaction starts out slowly (there is no HI present), but will speed up as the concentration of HI increases. Eventually both rates will level off (not always to the same level as shown by this example, however):



Chemists have found that there is a mathematical relationship that exists between the concentration of the reactants and products, once equilibrium has been reached, that is independent of the initial concentration of the participants. For any general reaction,

$$\mathrm{aA + bB \rightleftharpoons cD + dD} \tag{1.10.2.1}$$

an equilibrium constant expression can be written as:

$$\mathrm{K_{eq} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}} \tag{1.10.2.2}$$

This mathematical relationship exists for all equilibrium systems, and produces a constant ratio called the ***equilibrium constant, K_{eq}*** .

Law of Mass-Action

Equation 1.10.2.2is sometimes called the Law of Mass-Action.

This relationship will be very important to us for the next few units, so it is important that you understand how to set this relationship up and what it tells us about an equilibrium system.

The products of the reaction (C and D) are placed in the numerator, and their concentrations are raised to the power of the coefficients from the balanced equation. The reactants (A and B) are placed in the denominator, with their concentrations raised to the power of *their* coefficients.

Example 1.10.2.1

For the reaction between hydrogen and iodine gas to produce hydrogen iodide:

$$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)} \tag{1.10.2.3}$$

the equilibrium constant expression will be:

$$\mathrm{K_{eq}} = \frac{[\mathrm{HI}]^2}{[\mathrm{H_2}] \times [\mathrm{I_2}]} \tag{1.10.2.4}$$

Using the example we examined in our last section, equilibrium concentrations for each substance were measured at equilibrium and found to be:

At equilibrium:	$[\mathrm{H_2}] = 0.022\ \mathrm{M}$ $[\mathrm{I_2}] = 0.022\ \mathrm{M}$ $[\mathrm{HI}] = 0.156\ \mathrm{M}$

We substitute these values into our equilibrium expression and solve for $\mathrm{K_{eq}}$:

$$\mathrm{K_{eq}} = \frac{[\mathrm{HI}]^2}{[\mathrm{H_2}] \times [\mathrm{I_2}]} = \frac{(0.156)^2}{(0.022)(0.022)} = 50.3$$

The value of $\mathrm{K_{eq}}$, which has no units, is a constant for any particular reaction, and its value does not change unless the temperature of the system is changed. It does not depend on the initial concentrations used to reach the point of equilibrium.

For example, the following data were obtained for equilibrium concentrations of $\mathrm{H_2}$, $\mathrm{I_2}$ and HI, and the value of $\mathrm{K_{eq}}$ was calculated for each trial:

Trial	$[\mathrm{HI}]$	$[\mathrm{H_2}]$	$[\mathrm{I_2}]$	K_{eq}
1	0.156	0.0220	0.0220	50.3
2	0.750	0.106	0.106	50.1
3	1.00	0.820	0.0242	50.4
4	1.00	0.0242	0.820	50.4
5	1.56	0.220	0.220	50.3

Aside from accounting for slight experimental variation between trials, the value for $\mathrm{K_{eq}}$ is the same despite differences in equilibrium concentrations for the individual participants.

There is one other important point to make at this time.

K_{eq} relates the concentrations of products to reactants at equilibrium.

For aqueous solutions, concentration is often measured as mol · L^{-1}. For gases, concentration is often measured as partial pressure.

The concentrations of both aqueous solutions and gases change during the progress of a reaction. For reactions involving a solid or a liquid, while the ***amounts*** of the solid or liquid will change during a reaction, their ***concentrations*** (much like their densities) will ***not*** change during the reaction. Instead, their values will remain constant. Because they are constant, their values **are not included in the equilibrium constant expression**.

✓ Example 1.10.2.2

For example, consider the reaction showing the formation of solid calcium carbonate from solid calcium oxide and carbon dioxide gas:

$$\mathrm{CaO(s) + CO_2(g) \leftrightarrow CaCO_3(s)}$$

The equilibrium constant for this reaction is (before modification):

$$\mathrm{K_{eq} = \frac{[CaCO_3]}{[CaO] \times [CO_2]}}$$

But we remove those participants whose state is either a solid or a liquid, which leaves us with the following equilibrium constant expression:

$$\mathrm{K_{eq} = \frac{1}{[CO_2]}}$$

Summary

- There are a few possible ways a reaction can go: it can go to completion ($\text{reactants} \rightarrow \text{products}$), and it can occur but not go to completion. Instead, it will reach chemical equilibrium ($\text{reactants} \rightleftharpoons \text{products}$).
- Chemical equilibrium occurs when the number of particles becoming products is equal to the number of particles becoming reactants.
- A dynamic equilibrium is a state where the rate of the forward reaction is equal to the rate of the reverse reaction.

Vocabulary

- **Equilibrium** - A state that occurs when the rate of the forward reaction is equal to the rate of the reverse reaction.

1.10.3: The Equilibrium Constant - A Measure of How Far a Reaction Goes

Learning Objectives

- Write equilibrium constant expressions.
- Use equilibrium constant expressions to solve for unknown concentrations.
- Use known concentrations to solve for the equilibrium constants.
- Explain what the value of K means in terms of relative concentrations of reactants and products.

In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

The Equilibrium Constant

Equilibrium reactions are those that do not go to completion, but are in a state where the reactants are reacting to yield products and the products are reacting to produce reactants. In a reaction at equilibrium, the equilibrium concentrations of all reactants and products can be measured. The **equilibrium constant (K)** is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants. Sometimes, subscripts are added to the equilibrium constant symbol K, such as K_{eq}, K_{c}, K_{p}, K_{a}, K_{b}, and K_{sp}. These are all equilibrium constants and are subscripted to indicate special types of equilibrium reactions.

There are some rules about writing equilibrium constant expressions that need to be learned:

1. Concentrations of products are multiplied on the top of the expression. Concentrations of reactants are multiplied together on the bottom.
2. Coefficients in the equation become exponents in the equilibrium constant expression.
3. Solids, liquids, and solvents are assigned a value of 1, so their concentrations do not affect the value of K.

Example 1.10.3.1

Write the equilibrium constant expression for:

$$\mathrm{CO}\,(g)+3\mathrm{H_2}\,(g)\rightleftharpoons \mathrm{CH_4}\,(g)+\mathrm{H_2O}\,(g)$$

Solution

$$K=\frac{[\mathrm{CH_4}]\,[\mathrm{H_2O}]}{[\mathrm{CO}]\,[\mathrm{H_2}]^3}$$

*Note that the coefficients become exponents. Also, note that the concentrations of products in the numerator are *multiplied*. The same is true of the reactants in the denominator.

Example 1.10.3.2

Write the equilibrium constant expression for:

$$2\mathrm{TiCl_3}\,(s)+2\mathrm{HCl}\,(g)\rightleftharpoons 2\mathrm{TiCl_4}\,(s)+\mathrm{H_2}\,(g)$$

Solution

$$K=\frac{[\mathrm{H_2}]}{[\mathrm{HCl}]^2}$$

*Note that the solids have a value of 1, and multiplying or dividing by 1 does not change the value of K.

✓ Example 1.10.3.2

Write the equilibrium constant expression for:

$$\mathrm{P_4}\,(s) + 6\mathrm{Cl_2}\,(g) \rightleftharpoons 4\mathrm{PCl_3}\,(s)$$

Solution

$$K = \frac{1}{[\mathrm{Cl_2}]^6}$$

*Note that the only product is a solid, which is defined to have a value of 1. That leaves just 1 on top in the numerator.

✓ Example 1.10.3.3

Write the equilibrium constant expression for:

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

Solution

$$K = \left[\mathrm{H^+}\right]\left[\mathrm{OH^-}\right]$$

*Note that the water is the solvent, and thus has a value of 1. Dividing by 1 does not change the value of K.

Equilibrium Constant Expressions

The equilibrium constant value is the ratio of the concentrations of the products over the reactants. This means that we can use the value of K to predict whether there are more products or reactants at equilibrium for a given reaction. What can the value of K_{eq} tell us about a reaction?

- **If K_{eq} is very large**, the concentration of the products is much greater than the concentration of the reactants. The reaction essentially "goes to completion"; all, or most of, the reactants are used up to form the products.
- **If K_{eq} is very small**, the concentration of the reactants is much greater than the concentration of the products. The reaction does not occur to any great extent—most of the reactants remain unchanged, and there are few products produced.
- **When K_{eq} is not very large or very small** (close to a value of 1) then roughly equal amounts of reactants and products are present at equilibrium.

Here are some examples to consider:

Reaction	Chemical Equations	Equilibrium Constant
the decomposition of ozone, O_3	$2\,O_3(g) \rightleftharpoons 3\,O_2(g)$	$K_{eq} = 2.0 \times 10^{57}$
	K_{eq} is very large, indicating that mostly O_2 is present in an equilibrium system, with very little O_3.	
production of nitrogen monoxide	$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$	$K_{eq} = 1.0 \times 10^{-25}$
	Very little NO is produced by this reaction; N_2 and O_2 do not react readily to produce NO (this is lucky for us—otherwise we would have little oxygen to breath in our atmosphere!).	
reaction of carbon monoxide and water	$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$	$K_{eq} = 5.09$ (at 700 K)
	The concentrations of the reactants are very close to the concentrations of the products at equilibrium.	

If the equilibrium constant is 1 or nearly 1, it indicates that the molarities of the reactants and products are about the same. If the equilibrium constant value is a large number, like 100, or a very large number, like 1×10^{15}, it indicates that the products (numerator) are a great deal larger than the reactants. This means that at equilibrium, the great majority of the material is in the form of products and it is said that the "products are strongly favored". If the equilibrium constant is small, like 0.10, or very small, like 1×10^{-12}, it indicates that the reactants are much larger than the products and the reactants are strongly favored. With large K

values, most of the material at equilibrium is in the form of products and with small K values, most of the material at equilibrium is in the form of the reactants.

The equilibrium constant expression is an equation that we can use to solve for K or for the concentration of a reactant or product.

✓ Example 1.10.3.4

Determine the value of K for the reaction

$$\mathrm{SO_2}\,(g) + \mathrm{NO_2}\,(g) \rightleftharpoons \mathrm{SO_3}\,(g) + \mathrm{NO}\,(g)$$

when the equilibrium concentrations are: $[\mathrm{SO_2}] = 1.20\ \mathrm{M}$, $[\mathrm{NO_2}] = 0.60\ \mathrm{M}$, $[\mathrm{NO}] = 1.6\ \mathrm{M}$, and $\left[\mathrm{SO_3}\right] = 2.2\ \mathrm{M}$.

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{\left[\mathrm{SO_3}\right][\mathrm{NO}]}{[\mathrm{SO_2}][\mathrm{NO_2}]}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)(1.6)}{(1.20)(0.60)} = 4.9$$

✓ Example 1.10.3.5

Consider the following reaction:

$$\mathrm{CO}\,(g) + \mathrm{H_2O}\,(g) \rightleftharpoons \mathrm{H_2}\,(g) + \mathrm{CO_2}\,(g)$$

with $K = 1.34$. If the $[\mathrm{H_2O}] = 0.100\ \mathrm{M}$, $[\mathrm{H_2}] = 0.100\ \mathrm{M}$, and $[\mathrm{CO_2}] = 0.100\ \mathrm{M}$ at equilibrium, what is the equilibrium concentration of CO?

Solution

Step 1: Write the equilibrium constant expression:

$$K = \frac{[\mathrm{H_2}][\mathrm{CO_2}]}{[\mathrm{CO}][\mathrm{H_2O}]}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)(0.100)}{[\mathrm{CO}](0.100)}$$

Solving for $[\mathrm{CO}]$, we get: $[\mathrm{CO}] = 0.0746\ \mathrm{M}$

Summary

- The equilibrium constant expression is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants.
- If the value of K is greater than 1, the products in the reaction are favored. If the value of K is less than 1, the reactants in the reaction are favored. If K is equal to 1, neither reactants nor products are favored.

Vocabulary

- **Equilibrium constant** (K) - A mathematical ratio that shows the concentrations of the products divided by the concentrations of the reactants.

Contributions & Attributions

-
- Modified by Tom Neils (Grand Rapids Community College)

1.10.3.1: Calculating and Using Equilibrium Constants

Learning Objectives

- To understand how different phases affect equilibria.

When the products and reactants of an equilibrium reaction form a single phase, whether gas or liquid, the system is a homogeneous equilibrium. In such situations, the concentrations of the reactants and products can vary over a wide range. In contrast, a system whose reactants, products, or both are in more than one phase is a heterogeneous equilibrium, such as the reaction of a gas with a solid or liquid.

As noted in the previous section, the equilibrium constant expression is actually a ratio of activities. To simplify the calculations in general chemistry courses, the activity of each substance in the reaction is often approximated using a ratio of the molarity of a substance compared to the standard state of that substance. For substances that are liquids or solids, the standard state is just the concentration of the substance within the liquid or solid. Because the molar concentrations of pure liquids and solids normally do not vary greatly with temperature, the ratio of the molarity to the standard state for substances that are liquids or solids always has a value of 1. For example, for a compound such as CaF_2(s), the term going into the equilibrium expression is $[CaF_2]/[CaF_2]$ which cancels to unity. Thus, when the activities of the solids and liquids (including solvents) are incorporated into the equilibrium expression, they do not change the value.

Consider the following reaction, which is used in the final firing of some types of pottery to produce brilliant metallic glazes:

$$\mathrm{CO_2(g) + C(s) \rightleftharpoons 2\,CO(g)} \tag{1.10.3.1.1}$$

The glaze is created when metal oxides are reduced to metals by the product, carbon monoxide. The equilibrium constant expression for this reaction is as follows:

$$K = \frac{a^2_{\mathrm{CO}}}{a_{\mathrm{CO_2}}a_C} = \frac{[\mathrm{CO}]^2}{[\mathrm{CO_2}][1]} = \frac{[\mathrm{CO}]^2}{[\mathrm{CO_2}]} \tag{1.10.3.1.2}$$

The equilibrium constant for this reaction can also be written in terms of the partial pressures of the gases:

$$K_p = \frac{(P_{CO})^2}{P_{CO_2}} \tag{1.10.3.1.3}$$

Incorporating all the constant values into K' or K_p allows us to focus on the substances whose concentrations change during the reaction.

Although the activities of pure liquids or solids are not written explicitly in the equilibrium constant expression, these substances must be present in the reaction mixture for chemical equilibrium to occur. Whatever the concentrations of CO and $\mathrm{CO_2}$, the system described in Equation 1.10.3.1.1will reach chemical equilibrium only if a stoichiometric amount of solid carbon or excess solid carbon has been added so that some is still present once the system has reached equilibrium. As shown in Figure $1.10.3.1.1$, it does not matter whether 1 g or 100 g of solid carbon is present; in either case, the composition of the gaseous components of the system will be the same at equilibrium.

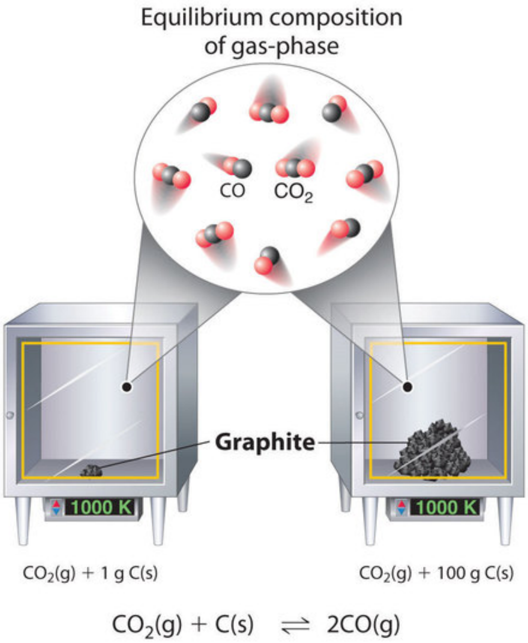


Figure 1.10.3.1.2: Effect of the Amount of Solid Present on Equilibrium in a Heterogeneous Solid–Gas System. In the system, the equilibrium composition of the gas phase at a given temperature, 1000 K in this case, is the same whether a small amount of solid carbon (left) or a large amount (right) is present.

✓ Example 1.10.3.1.1

Write each expression for K, incorporating all constants, and K_p for the following equilibrium reactions.

a. $\mathrm{PCl_3(l) + Cl_2(g) \rightleftharpoons PCl_5(s)}$
b. $\mathrm{Fe_3O_4(s) + 4\,H_2(g) \rightleftharpoons 3\,Fe(s) + 4\,H_2O(g)}$

Given: balanced equilibrium equations.

Asked for: expressions for K and K_p.

Strategy:

Find K by writing each equilibrium constant expression as the ratio of the concentrations of the products and reactants, each raised to its coefficient in the chemical equation. Then express K_p as the ratio of the partial pressures of the products and reactants, each also raised to its coefficient in the chemical equation.

Solution

This reaction contains a pure solid (PCl_5) and a pure liquid (PCl_3). Their activities are equal to 1, so when incorporated into the equilibrium constant expression, they do not change the value. So

$$K = \frac{1}{(1)[Cl_2]}$$

and

$$K_p = \frac{1}{(1)P_{Cl_2}}$$

This reaction contains two pure solids (Fe_3O_4 and Fe), which are each assigned a value of 1 in the equilibrium constant expressions:

$$K = \frac{(1)[H_2O]^4}{(1)[H_2]^4}$$

and

$$K_p = \frac{(1)(P_{H_2O})^4}{(1)(P_{H_2})^4}$$

Exercise 1.10.3.1.1

Write the expressions for K and K_p for the following reactions.

a. $\mathrm{CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)}$
b. $\underbrace{\mathrm{C_6H_{12}O_6}}_{glucose}\mathrm{(s) + 6\,O_2(g) \rightleftharpoons 6\,CO_2(g) + 6\,H_2O(g)}$

Answer a

$K = [\mathrm{CO_2}]$ and $K_p = P_{\mathrm{CO_2}}$

Answer b

$$K = \frac{[CO_2]^6[H_2O]^6}{[O_2]^6} \text{ and } K_p = \frac{(P_{CO_2})^6(P_{H_2O})^6}{(P_{O_2})^6}$$

For reactions carried out in solution, the solvent is assumed to be pure, and therefore is assigned an activity equal to 1 in the equilibrium constant expression. The activities of the solutes are approximated by their molarities. The result is that the equilibrium constant expressions appear to only depend upon the concentrations of the solutes.

> *The activities of pure solids, pure liquids, and solvents are defined as having a value of '1'. Often, it is said that these activities are "left out" of equilibrium constant expressions. This is an unfortunate use of words. The activities are not "left out" of equilibrium constant expressions. Rather, because they have a value of '1', they do not change the value of the equilibrium constant when they are multiplied together with the other terms. The activities of the solutes are approximated by their molarities.*

Summary

An equilibrated system that contains products and reactants in a single phase is a homogeneous equilibrium; a system whose reactants, products, or both are in more than one phase is a heterogeneous equilibrium.

Contributors and Attributions

- Anonymous
- Modified by Tom Neils (Grand Rapids Community College)

1.10.3.2: Practice Equilibrium Constant Expressions

Exercise 1.10.3.2.1

Which of the following is true at equilibrium conditions? (Choose all that apply.)

a. The forward reaction rate is zero.

b. The amount of reactants left is zero.

c. The forward reaction rate is equal to the reverse reaction rate.

d. The amount of products is equal to the amount of reactants.

e. The amount of products that is present is constant.

f. The product stops forming.

Answer

c, e.

Exercise 1.10.3.2.1

What does it mean if a reversible reaction has an equilibiurm constant (K_{eq}) with a very small value?

a. The reaction rate is very slow.

b. There are more reactants than products at equilibrium.

c. There are more products than reactants at equilibrium.

d. The forward reaction is slower than the reverse reaction.

Answer

b

Exercise 1.10.3.2.1

Which of the following would be defined as having a value of 1 in an equilibrium constant expression? (Choose all that apply.)

a. a solid

b. a gas

c. a liquid

d. a solute dissolved in a solution

Answer

a and c.

Exercise 1.10.3.2.1

$$2SO_{2(g)} + O_{2(g)} \longleftrightarrow 2SO_{3(g)} \tag{1.10.3.2.1}$$

What is the equilibrium constant expression of the reaction above?

Answer

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

1.10.4: Calculations with Equilibrium Constants

Iron is an important component of red blood cells. Patients who have low iron will usually be anemic and have a lower than normal number of red blood cells. One way to assess serum iron concentration is with the use of Ferrozine, a complex organic molecule. Ferrozine forms a product with Fe^{3+}, producing a pink color. In order to determine factors affecting the reaction, we need to measure the equilibrium constant. If the equilibrium does not lie far in the direction of products, precautions need to be taken when using this material to measure iron in serum.

Calculations with Equilibrium Constants

The general value of the equilibrium constant gives us information about whether the reactants or the products are favored at equilibrium. Since the product concentrations are in the numerator of the equilibrium expression, a $K_{\mathrm{eq}} > 1$ means that the products are favored over the reactants. A $K_{\mathrm{eq}} < 1$ means that the reactants are favored over the products.

Though it would often seem that the K_{eq} value would have various units depending on the values of the exponents in the expression, the general rule is that any units are dropped. All K_{eq} values will be reported as having no units.

Example 1.10.4.1

Equilibrium occurs when nitrogen monoxide gas reacts with oxygen gas to form nitrogen dioxide gas:

$$2\mathrm{NO}\,(g) + \mathrm{O}_2\,(g) \rightleftharpoons 2\mathrm{NO}_2\,(g)$$

At equilibrium at $230^{\mathrm{o}}\mathrm{C}$, the concentrations are measured to be $[\mathrm{NO}] = 0.0542\ \mathrm{M}$, $[\mathrm{O}_2] = 0.127\ \mathrm{M}$, and $[\mathrm{NO}_2] = 15.5\ \mathrm{M}$. Calculate the equilibrium constant at this temperature.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $[\mathrm{NO}] = 0.0542\ \mathrm{M}$
- $[\mathrm{O}_2] = 0.127\ \mathrm{M}$
- $[\mathrm{NO}_2] = 15.5\ \mathrm{M}$

Unknown

The equilibrium expression is first written according to the general form in the text. The equilibrium values are substituted into the expression and the value calculated.

Step 2: Solve.

$$K_{\mathrm{eq}} = \frac{[\mathrm{NO}_2]^2}{[\mathrm{NO}]^2\,[\mathrm{O}_2]}$$

Substituting in the concentrations at equilibrium:

$$K_{\mathrm{eq}} = \frac{(15.5)^2}{(0.0542)^2\,(0.127)} = 6.44 \times 10^5$$

Step 3: Think about your result.

The equilibrium concentration of the product NO_2 is significantly higher than the concentrations of the reactants NO and O_2. As a result, the K_{eq} value is much larger than 1, an indication that the product is favored at equilibrium.

The equilibrium expression only shows those substances whose concentrations are variable during the reaction. A pure solid or a pure liquid does not have a concentration that will vary during a reaction. Therefore, an equilibrium expression omits pure solids and liquids, and only shows the concentrations of gases and aqueous solutions. The decomposition of mercury (II) oxide can be shown by the following equation, followed by its equilibrium expression.

$$2\mathrm{HgO}\,(s) \rightleftharpoons 2\mathrm{Hg}\,(l) + \mathrm{O}_2\,(g) \qquad K_{\mathrm{eq}} = [\mathrm{O}_2]$$

The stoichiometry of an equation can also be used in a calculation of an equilibrium constant. At 40°C, solid ammonium carbamate decomposes to ammonia and carbon dioxide gases.

$$\text{NH}_4\text{CO}_2\text{NH}_2\,(s) \rightleftharpoons 2\text{NH}_3\,(g) + \text{CO}_2\,(g)$$

At equilibrium, the $[\text{CO}_2]$ is found to be $4.71 \times 10^{-3}\ \text{M}$. Can the K_{eq} value be calculated from that information alone? Because the ammonium carbamate is a solid, it is not present in the equilibrium expression.

$$K_{\text{eq}} = [\text{NH}_3]^2\,[\text{CO}_2]$$

The stoichiometry of the chemical equation indicates that as the ammonium carbamate decomposes, $2\ \text{mol}$ of ammonia gas is produced for every $1\ \text{mol}$ of carbon dioxide. Therefore, at equilibrium, the concentration of the ammonia will be twice the concentration of carbon dioxide. So,

$$[\text{NH}_3] = 2 \times (4.71 \times 10^{-3}) = 9.42 \times 10^{-3}\ \text{M}$$

Substituting these values into the K_{eq} expression:

$$K_{\text{eq}} = (9.42 \times 10^{-3})^2\,(4.71 \times 10^{-3}) = 4.18 \times 10^{-7}$$

Using Equilibrium Constants

The equilibrium constants are known for a great many reactions. Hydrogen and bromine gases combine to form hydrogen bromide gas. At 730°C, the equation and K_{eq} are given below.

$$\text{H}_2\,(g) + \text{Br}_2\,(g) \rightleftharpoons 2\text{HBr}\,(g) \qquad K_{\text{eq}} = 2.18 \times 10^6$$

A certain reaction is begun with only HBr. When the reaction mixture reaches equilibrium at 730°C, the concentration of bromine gas is measured to be $0.00243\ \text{M}$ What is the concentration of the H_2 and the HBr at equilibrium?

Since the reaction begins with only HBr and the mole ratio of H_2 to Br_2 is 1:1, the concentration of H_2 at equilibrium is also $0.00243\ \text{M}$ The equilibrium expression can be rearranged to solve for the concentration of HBr at equilibrium:

$$\begin{aligned} K_{\text{eq}} &= \frac{[\text{HBr}]^2}{[\text{H}_2]\,[\text{Br}_2]} \\ [\text{HBr}] &= \sqrt{K_{\text{eq}}\,[\text{H}_2]\,[\text{Br}_2]} \\ &= \sqrt{2.18 \times 10^6\,(0.00243)\,(0.00243)} = 3.59\ \text{M} \end{aligned}$$

Since the value of the equilibrium constant is very high, the concentration of HBr is much greater than that of H_2 and Br_2 at equilibrium.

Summary

1.10.4.1: Practice Equilibrium Constants

Exercise 1.10.4.1.1

$$3H_{2(g)} + N_{2(g)} \longleftrightarrow 2NH_{3(g)} \tag{1.10.4.1.1}$$

What is the equilibrium constant expression of the reaction above?

Answer

$$K_{eq} = \frac{[NH_3]^2}{[H_2]^3[N_2]}$$

Exercise 1.10.4.1.1

$$2SO_{2(g)} + O_{2(g)} \longleftrightarrow 2SO_{3(g)} \tag{1.10.4.1.2}$$

The K_{eq} of the reaction above is 3.4 x 10^2. Is the equilibrium mixture mostly SO_2 and O_2 or mostly SO_3?

Answer

mostly SO_3

Exercise 1.10.4.1.1

$$C_{(s)} + 2S_{(g)} \longleftrightarrow CS_{2(g)} \tag{1.10.4.1.3}$$

What is the equilibrium constant expression of the reaction above?

Answer

$$K_{eq} = \frac{[CS_2]}{[S]^2}$$

Exercise 1.10.4.1.1

$$Pb^{2+}_{(aq)} + 2Cl^{1-}_{(aq)} \longleftrightarrow PbCl_{2(s)} \tag{1.10.4.1.4}$$

What is the equilibrium constant expression of the reaction above?

Answer

$$K_{eq} = \frac{1}{[Pb^{2+}][Cl^{1-}]^2}$$

1.10.5: Disturbing a Reaction at Equilibrium- Le Châtelier’s Principle

Learning Objectives

- State Le Chatelier's Principle.
- Describe the effect of concentration on an equilibrium system.
- Describe the effect of temperature as a stress on an equilibrium system.

When a reaction has reached equilibrium with a given set of conditions, if the conditions are not changed, the reaction will remain at equilibrium forever. The forward and reverse reactions continue at the same equal and opposite rates and the macroscopic properties remain constant.

It is possible, however, to disturb that equilibrium by changing conditions. For example, you could increase the concentration of one of the products, or decrease the concentration of one of the reactants, or change the temperature. When a change of this type is made within a reaction at equilibrium, the reaction is no longer in equilibrium. When you alter something in a reaction at equilibrium, chemists say that you put **stress** on the equilibrium. When this occurs, the reaction will no longer be in equilibrium and the reaction itself will begin changing the concentrations of reactants and products until the reaction comes to a new position of equilibrium. How a reaction will change when a stress is applied can be explained and predicted. That is the topic of this section.

Le Chatelier's Principle

Think back to our escalator example, with you walking up a downward moving escalator. With the rate of the moving stairs and your walking evenly matched, you appear to be at a standstill. But what happens if the escalator begins moving just a little faster? If you want to maintain the same position you had, at some specific point between the bottom and the top of the stairs, you'll also need to make some adjustments. Chemical systems at equilibrium tend to make these adjustments as well.

In the late 1800's, a chemist by the name of Henry-Louis Le Chatelier was studying stresses that were applied to chemical equilibria. He formulated a principle from this research and, of course, the principle is called Le Chatelier's Principle. **Le Chatelier's Principle** states that when a stress is applied to a system at equilibrium, the equilibrium will shift in a direction to partially counteract the stress and once again reach equilibrium. Le Chatelier's principle is not an explanation of what happens on the molecular level to cause the equilibrium shift, it is simply a quick way to determine which way the reaction will run in response to a stress applied to the system at equilibrium.

Le Chatelier's Principle

If a system at **equilibrium** is subjected to an external stress, the equilibrium will shift to **minimize** the effects of that stress.

Equilibrium is all about rates—the rate of the forward reaction is equal to the rate of the reverse reaction. External stresses are factors that will cause the rate of either the forward or reverse reaction to change, throwing the system out of balance. Le Chatelier's Principle allows us to predict how this will affect our system.

In our unit on Kinetics we examined factors that influenced reaction rates. Recall these factors:

1. concentration
2. pressure and volume
3. temperature
4. catalysts

We will see how changing these factors affects a system at equilibrium.

Effect of Concentration Changes on a System at Equilibrium

For instance, if a stress is applied by increasing the concentration of a reactant, the reaction will adjust in such a way that the reactants and products can get back to equilibrium. In the case of too much reactant, the reaction will use up some of the reactant to make more product. It is said in this scenario that the reaction "shifts to the products" or "shifts to the right". If the concentration of a product is increased, there is an opposite effect. The reaction will use up some of the product to make more reactant. The reaction "shifts to the reactants" or "shifts to the left".

What if some reactant or product is removed? If a stress is applied by lowering a reactant concentration, the reaction will try to replace some of the missing reactant. It uses up some of the product to make more reactant, and the reaction "shifts to the reactants". If a stress is applied by reducing the concentration of a product, the equilibrium position will shift toward the products.

1.10.5.1: The Effect of a Concentration Change on Equilibrium

Consider the following system under equilibrium:

$$\underbrace{\mathrm{Fe^{3+}(aq)}}_{\text{colorless}} + \underbrace{\mathrm{SCN^{-}(aq)}}_{\text{colorless}} \rightleftharpoons \underbrace{\mathrm{FeSCN^{2+}(aq)}}_{\text{red}}$$

If more Fe^{3+} is added to the reaction, what will happen?

According to Le Chatelier's Principle, the system will react to minimize the stress. Since $\mathrm{Fe^{3+}}$ is on the reactant side of this reaction, the rate of the forward reaction will increase in order to "use up" the additional reactant. This will cause the equilibrium to **shift to the right**, producing more $\mathrm{FeSCN^{2+}}$. For this particular reaction, we will be able to see that this has happened, as the solution will become a darker red color.

There are a few different ways to state what happens here when more $\mathrm{Fe^{3+}}$ is added, all of which have the same meaning:

- equilibrium shifts to the right
- equilibrium shifts to the product side
- the forward reaction is favored

What changes does this cause in the concentrations of the reaction participants?

Changes within reaction participants

$\mathrm{Fe^{3+}}$	Since this is what was added to cause the stress, the concentration of $\mathrm{Fe^{3+}}$ will increase. (A shorthand way to indicate this: $[\mathrm{Fe}]^{3+}\uparrow$ (Reminder: the square brackets represent "concentration")
$\mathrm{SCN^{-}(aq)}$	Equilibrium will shift to the right, which will use up the reactants. The concentration of $\mathrm{SCN^{-}(aq)}$ will decrease $[\mathrm{SCN}]^{-}\downarrow$ as the rate of the forward reaction increases.
$\mathrm{FeSCN^{2+}}$	When the forward reaction rate increases, more products are produced, and the concentration of $\mathrm{FeSCN^{2+}}$ will increase. $[\mathrm{FeSCN}]^{2+}\uparrow$

How about the value of K_{eq}? Notice that the concentration of some reaction participants have increased, while others have decreased. Once equilibrium has re-established itself, the value of K_{eq} will be unchanged.

> *The value of K_{eq} does not change when changes in concentration cause a shift in equilibrium.*

What if more $\mathrm{FeSCN^{2+}}$ is added?

Again, equilibrium will shift to use up the added substance. In this case, equilibrium will shift to favor the *reverse* reaction, since the reverse reaction will use up the additional $\mathrm{FeSCN^{2+}}$.

- equilibrium shifts to the left
- equilibrium shifts to the reactant side
- the reverse reaction is favored

How do the concentrations of reaction participants change?

Change of concentrations of reaction participants when adding substance

$\mathrm{Fe^{3+}}$	$[\mathrm{Fe}]^{3+}\uparrow$ as the reverse reaction is favored
$\mathrm{SCN^{-}(aq)}$	$[\mathrm{SCN}]^{-}\uparrow$ as the reverse reaction is favored
$\mathrm{FeSCN^{2+}}$	$[\mathrm{FeSCN}]^{2+}\uparrow$ because this is the substance that was added

Concentration can also be changed by **removing** a substance from the reaction. This is often accomplished by adding another substance that reacts (in a side reaction) with something already in the reaction.

Let's remove SCN⁻ from the system (perhaps by adding some Pb^{2+} ions—the lead(II) ions will form a precipitate with SCN^-, removing them from the solution). What will happen now? Equilibrium will shift to **replace** SCN^-—the reverse reaction will be favored because that is the direction that produces more SCN^-.

- equilibrium shifts to the left
- equilibrium shifts to the reactant side
- the reverse reaction is favored

How do the concentrations of reaction participants change?

Change of concentrations of reaction participants when removing a substance

Fe^{3+}	$[Fe]^{3+}$ ↑ as the reverse reaction is favored
SCN^-	$[SCN]^-$ ↑ as the reverse reaction is favored (but also ↓ because it was removed)
$FeSCN^{2+}$	$[FeSCN]^{2+}$ ↑ because this is the substance that was added

---

1.10.5.2: The Effect of a Volume Change on Equilibrium

Changing the pressure or volume of a container enclosing an equilibrium system will only affect the reaction if gases are present.

You may remember from earlier chemistry classes that equal volumes of gases contain an equal number of particles and, under standard conditions of temperature and pressure (STP), one mole of gas occupies a volume of 22.4 L. This is known as the molar volume of gases. So, two moles of any gas will occupy a volume of 44.8 L and one-half mole would occupy 11.2 L.

How does changing pressure and volume affect equilibrium systems?

- If you increase the pressure of a system at equilibrium (typically by reducing the volume of the container), the stress will best be reduced by reaction that favors the side with the ***fewest*** moles of gas, since fewer moles will occupy the smallest volume.
- Conversely, if you decrease the pressure (by increasing the volume of the container), equilibrium will shift to favor the side with the ***most*** moles of gas, since more moles will occupy a greater volume.
- If both sides of the equation have the same number of moles of gas, then there will be no change in the position of equilibrium.

When considering the effect of changing volume or pressure on equilibrium systems, be sure to only count the number of moles of **gases** on each side of the equation. Solids, liquids, and aqueous solutions will not be affected by changing pressure and volume.

✓ Example 1.10.5.2.1

Predict the effect on equilibrium when the pressure is increased for the following reaction:

$$\mathrm{N_2O_4(g) \rightleftharpoons 2\,NO_2(g)}$$

Solution

The reactant side of the equation has 1 mole of a gas; the product side has 2 moles. Increasing the pressure favors the side with the fewest moles of gas, so the equilibrium will shift to the left (the reverse reaction will be favored).

1.10.5.3: The Effect of Temperature Changes on Equilibrium

Learning Objectives

- Explain how temperature changes affect a system at equilibrium.

When temperature is the stress that affects a system at equilibrium, there are two important consequences:

- an increase in temperature will favor that reaction direction that ***absorbs*** heat (i.e. the endothermic reaction)
- the value of K_{eq} will change

Consider the following equilibrium system:

$$N_2O_4(g) \leftrightarrow 2\,NO_2(g)$$

$$\text{with } \Delta H^\circ = 58 \cdot 0 \text{ kJ}$$

We see by the sign of ΔH° that the forward reaction is endothermic. Heat is absorbed (required as a reactant) when the reaction proceeds as

$$N_2O_4(g) \rightarrow 2\,NO_2(g)$$

By adding more heat, equilibrium will shift to use up the additional heat, thus favoring this forward direction.

Why will K_{eq} change, when it did not change when concentration, pressure, and volume were the applied stresses?

When temperature changes cause an equilibrium to shift, one entire side of the reaction equation is favored over the other side. Mathematically, this will alter the value of K_{eq} as follows:

$$K_{eq} = \frac{[\text{products}]}{[\text{reactants}]}$$

\ce{K_{eq}}

forward and reverse reactions

if the forward reaction is favored	more products are produced; fewer reactants	K_{eq} will increase
if the reverse reaction is favored	fewer products; more reactants	K_{eq} will decrease

So in our example given above, increasing the temperature will favor the forward direction. The value of K_{eq} will increase. **Removing heat (making the system colder) will favor the exothermic reaction**—the exothermic reaction ***releases heat*** to the surroundings, thus "replacing" the heat that has been removed.

1.10.5.4: The Path of a Reaction and the Effect of a Catalyst

The addition of a catalyst to an equilibrium system is a final stress factor. We consider how adding a catalyst affects the following:

$$\mathrm{N_2(g) + O_2(g) \leftrightharpoons 2\,NO(g)}$$

Adding a catalyst to this, or any other equilibrium system, will ***not*** affect the position of an equilibrium. A catalyst speeds up both the forward and the reverse reactions, so there is no uneven change in reaction rates. Generally, a catalyst will help a reaction to reach the point of equilibrium *sooner*, but it will not affect the equilibrium otherwise.

1.10.5.5: Practice Le Chatelier's Principle

Exercise 1.10.5.5.1

$$3H_{2(g)} + N_{2(g)} \longleftrightarrow 2NH_{3(g)} \text{ with} \Delta H = -92kJ \tag{1.10.5.5.1}$$

Determine whether each of the changes below would cause the equilibrium above to shift left, shift right, or have no shift.

a. add N_2

b. increase temperature

c. add catalyst

d. remove H_2

e. increase volume

Answer

a. shift right; b. shift left; c. no shift; d. shift left; e. shift left

Exercise 1.10.5.5.1

$$2SO_{2(g)} + O_{2(g)} \longleftrightarrow 2SO_{3(g)} + heat \tag{1.10.5.5.2}$$

Determine whether each of the changes below would cause the equilibrium above to shift left, shift right, or have no shift.

a. add O_2

b. increase temperature

c. add catalyst

d. remove SO_3

e. increase volume

Answer

a. shift right; b. shift left ; c. no shift; d. shift right; e. shift left

Exercise 1.10.5.5.1

$$N_{2(g)} + O_{2(g)} + heat \longleftrightarrow 2NO_{(g)} \tag{1.10.5.5.3}$$

Determine whether each of the changes below would cause the equilibrium above to shift left, shift right, or have no shift.

a. remove O_2

b. increase temperature

c. add catalyst

d. add NO

e. increase volume

Answer

a. shift left; b. shift right; c. no shift; d. shift left; e. no shift

SECTION OVERVIEW

1.11: Acids and Bases

1.11.1: Acids- Properties and Examples

Learning Objectives

- Examine properties of acids.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{1.11.1.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

It should not be hard for you to name several common acids (but you might find that listing bases is a little more difficult). Below is a partial list of some common acids, along with some chemical formulas:

Table 1.11.1.1: Common Acids and Their Uses

Chemist Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) and stomach acid is HCl	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.
sulfuric acid, H_2SO_4		Used in car batteries, and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fizz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $C_6H_8O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$	aspirin	The active ingredient in aspirin.

What exactly makes an acid an acid, and what makes a base act as a base? Take a look at the formulas given in the above table and take a guess.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. A colorless, highly pungent solution of hydrogen chloride (HCl) in water. Hydrochloric acid is usually prepared by treating HCl with water.

$$\mathrm{HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)}$$

Hydrochloric acid can therefore be used to prepare chloride salts. Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases.

Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula $\mathrm{H_2SO_4}$. Sulfuric acid is a diprotic acid and has a wide range of applications including use in domestic acidic drain cleaners,[as an electrolyte in lead-acid batteries, and in various cleaning agents. It is also a central substance in the chemical industry.

Figure 1.11.1.1: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating property, removing water ($\mathrm{H_2O}$) from other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 1.11.1.2).

$$\mathrm{M + H_2SO_4 \rightarrow M(SO_4) + H_2} \tag{1.11.1.2}$$

Nitric Acid

Nitric acid ($\mathrm{HNO_3}$) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide ($\mathrm{NO_2(g)}$) with water.

$$\mathrm{3\,NO_2(g) + H_2O(l) \rightarrow 2\,HNO_3(ag) + NO(g)}$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals (e.g., nitric acid with magnesium, manganese or zinc will liberate $\mathrm{H_2}$ gas):

$$\mathrm{Mg + 2\,HNO_3 \rightarrow Mg(NO_3)_2 + H_2}$$

$$\mathrm{Mn + 2\,HNO_3 \rightarrow Mn(NO_3)_2 + H_2}$$

$$\mathrm{Zn + 2\,HNO_3 \rightarrow Zn(NO_3)_2 + H_2}$$

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard it poses is chemical burn, as it carries out acid hydrolysis with proteins (amide) and fats (ester) which consequently decomposes living tissue (Figure 1.11.1.2). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin

Figure 1.11.1.2: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula H_2CO_3 and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of $H_2CO_3(aq)$. Carbonic acid, which is a weak acid, forms two kinds of salts: the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve, producing calcium bicarbonate—which leads to many limestone features such as stalactites and stalagmites. Carbonic acid is a polyprotic acid, specifically it is diprotic, meaning that it has two protons which may dissociate from the parent molecule.

When carbon dioxide dissolves in water, it exists in chemical equilibrium (discussed in Chapter 15), producing carbonic acid:

$$CO_2 + H_2O \rightleftharpoons H_2CO_3$$

The reaction can be pushed to favor the reactants to generate $CO_2(g)$ from solution, which is key to the bubbles observed in carbonated beverages (Figure 1.11.1.3).

Figure 1.11.1.3: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

Formic Acid

Formic acid (HCO_2H) is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally, most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

Citric Acid

Citric acid ($C_6H_8O_7$) is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the TCA cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is used as a flavoring and preservative in food and beverages, especially soft drinks.

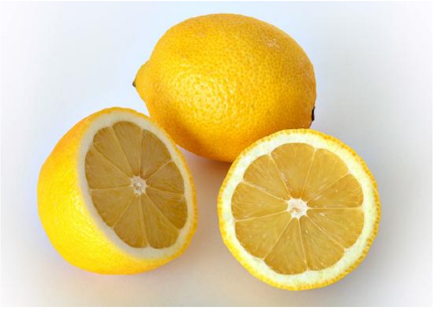


Figure 1.11.1.4: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years.

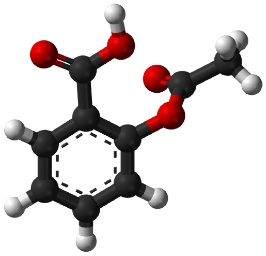


Figure 1.11.1.5: Ball-and-stick model of the aspirin molecule. (Public Domain; Ben Mills).

Aspirin is a white, crystalline, weakly acidic substance.

Summary

A brief summary of key aspects of several acids commonly encountered by students was given. Acids are a distinct class of compounds because of the properties of their aqueous solutions.

Contributions & Attributions

-
- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

1.11.1.1: Bases- Properties and Examples

Learning Objectives

- Examine properties of bases.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base, while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.

Figure 1.11.1.1.1: Phenolphthalein indicator in presence of base.

Warning!

Tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words—don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate. Various common bases and corresponding uses are given in Table 1.11.1.1.2

Table 1.11.1.1.1: Common Bases and Corresponding Uses

Some Common Bases	Uses
sodium hydroxide, NaOH (lye or caustic soda)	Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners.
potassium hydroxide, KOH (lye or caustic potash)	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$ (slaked lime)	Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil.
aluminum hydroxide	Used in water purification and as an ingredient in antacids.
ammonia, NH_3	Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula NaOH. It is a white solid ionic compound consisting of sodium cations $\mathrm{Na^+}$ and hydroxide anions $\mathrm{OH^-}$.

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction:

$$\mathrm{NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)}$$

The resulting solution is usually colorless and odorless and feels slippery when it comes in contact with skin.

Figure 1.11.1.1.1: Sample of sodium hydroxide as pellets in a watch glass. (Public Domain; Walkerma.)

Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula KOH, and is commonly called caustic potash. Along with sodium hydroxide (NaOH), this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its corrosive nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathrm{KOH(s) \rightarrow K^+(aq) + OH^-(aq)}$$

Concentrated aqueous solutions are sometimes called *potassium lyes*.

Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $\mathrm{Mg(OH)_2}$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.

Figure 1.11.1.1.1: Bottle of Antacid tablets. (CC BY 2.,5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $\mathrm{Mg(OH)_2}$. However, a weak concentration of dissociated ions can be found in solution:

$$\mathrm{Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\,OH^-(aq)}$$

Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathrm{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$\mathrm{Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\,OH^-(aq)}$$

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathrm{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of "smelling salts," and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

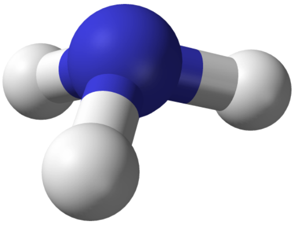


Figure 1.11.1.1.1: Ball-and-stick model of the ammonia molecule. (Public Domain; Ben Mills).

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from $\mathrm{H_2O}$ to yield ammonium and hydroxide ions:

$$\mathrm{NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)}$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

- A brief summary of properties of bases was given.
- The properties of bases mostly contrast those of acids.
- Bases have many, varied uses.

1.11.2: Molecular Definitions of Acids and Bases

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major classifications of substances known as acids or bases. The theory developed by Svante Arrhenius in 1883, the Arrhenius definition, states that an acid produces H^+ in solution and a base produces OH^-. Later, two more sophisticated and general theories were proposed. These theories are the Brønsted-Lowry and Lewis definitions of acids and bases. This section will cover the Arrhenius and Brønsted-Lowry theories; the Lewis theory is discussed elsewhere.

The Arrhenius Theory of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. An **Arrhenius acid** is a compound that **increases the concentration of H^+ ions** that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$\mathrm{HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)}$$

In this reaction, hydrochloric acid (HCl) dissociates completely into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$\mathrm{HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)}$$

An **Arrhenius base** is a compound that **increases the concentration of OH^- ions** that are present when added to water. The dissociation is represented by the following equation:

$$\mathrm{NaOH\ (aq) \rightarrow Na^+\ (aq) + OH^-\ (aq)}$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium (Na^+) and hydroxide (OH^-) ions when dissolved in water, thereby releasing OH^- ions into solution.

> *Arrhenius acids are substances which produce hydrogen ions in solution and Arrhenius bases are substances which produce hydroxide ions in solution.*

Limitations to the Arrhenius Theory

The Arrhenius theory has many more limitations than the other two theories. The theory does not explain the weak base ammonia (NH_3), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself. The Arrhenius definition of acid and base is also limited to aqueous (i.e., water) solutions.

The Brønsted-Lowry Theory of Acids and Bases

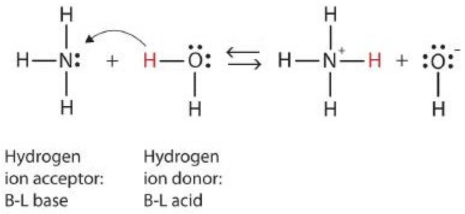
In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, **a Brønsted-Lowry acid** is a **proton donor (PD)**, while **a Brønsted-Lowry base** is a **proton acceptor (PA)**.

> *A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.*

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$\mathrm{NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)} \tag{1.11.2.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:

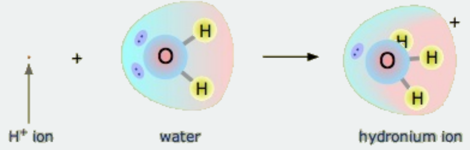


Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

The Hydronium Ion

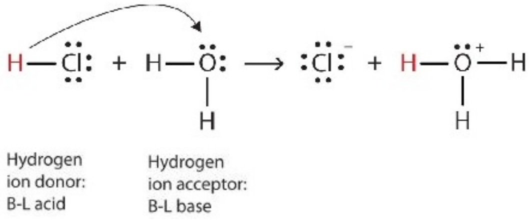
A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+ to represent the hydronium ion.



With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$\mathrm{HCl(g)} + \mathrm{H_2O}(\ell) \rightarrow \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)} \tag{1.11.2.2}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense, but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.
- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.

- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

✓ Example 1.11.2.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule, just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$\mathrm{C_6H_5NH_2(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)}$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

? Exercise 1.11.2.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$\mathrm{H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2\,-} + H_3O^+}$$

Answer

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: H_2O

? Exercise 1.11.2.2

Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. HPO_4^{2-}
c. H_3PO_4
d. NH_4^+
e. $CH_3NH_3^+$

Answer

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H^+. This eliminates HCl, $\mathrm{H_3PO_4}$, $\mathrm{NH_4^+}$ and $\mathrm{CH_3NH_3^+}$ because they are Bronsted-Lowry acids. They all give away protons. In the case of $\mathrm{HPO_4^{2\,-}}$, consider the following equation:

$$\mathrm{HPO_4^{2\,-}(aq) + H_2O(l) \rightarrow PO_4^{3\,-}(aq) + H_3O^+(aq)}$$

Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-}. Now consider the following equation:

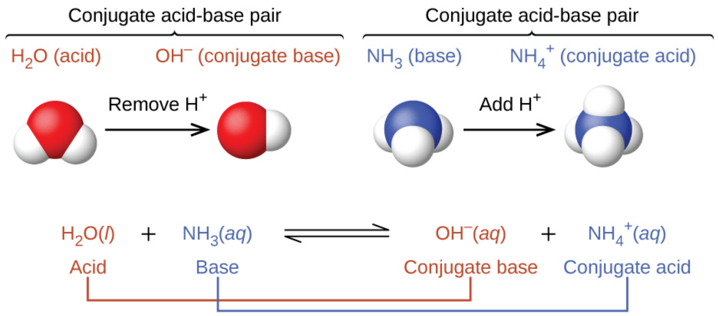
$$\mathrm{HPO_4^{2\,-}(aq) + H_2O(l) \rightarrow H_2PO_4^- + OH^-(aq)}$$

In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $H_2PO_4^-$ and OH^-. Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) HPO_4^{2-}.**

Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, $\mathrm{OH^-}$, and the conjugate acid of ammonia, $\mathrm{NH_4^+}$:



This figure has three parts in two rows. In the first row, two diagrams of acid-base pairs are shown. On the left, a space filling model of H subscript 2 O is shown with a red O atom at the center and two smaller white H atoms attached in a bent shape. Above this model is the label "H subscript 2 O (acid)" in purple. An arrow points right, which is labeled "Remove H superscript plus." To the right is another space filling model with a single red O atom to which a single smaller white H atom is attached. The label in purple above this model reads, "O H superscript negative (conjugate base)." Above both of these red and white models is an upward pointing bracket that is labeled "Conjugate acid-base pair." To the right is a space filling model with a central blue N atom to which three smaller white H atoms are attached in a triangular pyramid arrangement. A label in green above reads "N H subscript 3 (base)." An arrow labeled "Add H superscript plus" points right. To the right of the arrow is another space filling model with a blue central N atom and four smaller white H atoms in a tetrahedral arrangement. The green label above reads "N H subscript 3 superscript plus (conjugate acid)." Above both of these blue and white models is an upward pointing bracket that is labeled "Conjugate acid-base pair." The second row of the figure shows the chemical reaction, H subscript 2 O (l) is shown in purple, and is labeled below in purple as "acid," plus N H subscript 3 (a q) in green, labeled below in green as "base," followed by a double sided arrow arrow and O H superscript negative (a q) in purple, labeled in purple as "conjugate base," plus N H subscript 4 superscript plus (a q)" in green, which is labeled in green as "conjugate acid." The acid on the left side of the equation is connected to the conjugate base on the right with a purple line. Similarly, the base on the left is connected to the conjugate acid on the right side.

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-.

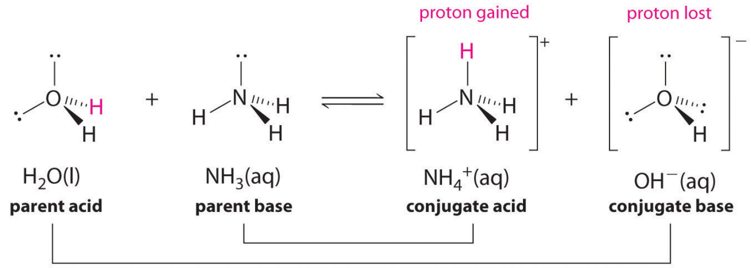


Figure 1.11.2.1. The pairing of parent acids and bases with conjugate acids and bases.

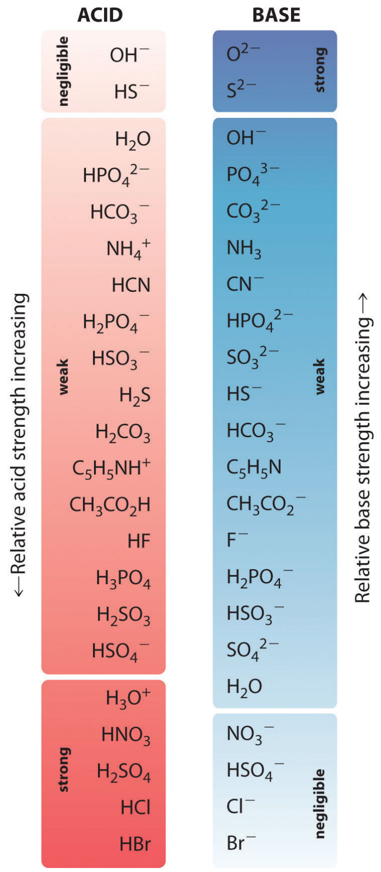


Figure 1.11.2.1: The Relative Strengths of Some Common Conjugate Acid-Base Pairs. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

✓ Example 1.11.2.2

Identify the conjugate acid-base pairs in this equilibrium.

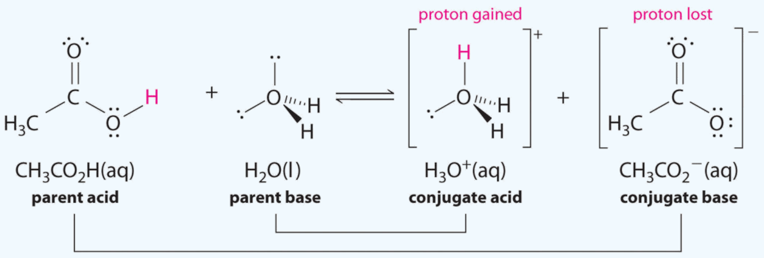
$$CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-$$

Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid-base pairs:

- the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and
- the parent base and its conjugate acid (H_3O^+/H_2O).



✓ Example 1.11.2.3

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H_2O and OH^-, where H_2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base.

The other pair consists of $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

? Exercise 1.11.2.3

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer

H_2O (acid) and OH^- (base); NH_2^- (base) and NH_3 (acid)

1.11.3: Strong and Weak Acids and Bases

Learning Objectives

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.
- Determine if a salt produces an acidic or a basic solution.

Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider HCl(aq). When HCl is dissolved in H_2O, it completely dissociates into H^+(aq) and Cl^-(aq) ions; all the HCl molecules become ions:

$$HCl \overset{100\%}{\rightarrow} H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2 \xrightarrow{\sim 5\%} H^+(aq) + C_2H_3O_2^-(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as **a reversible reaction**:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

As it turns out, there are very few strong acids, which are given in Table 1.11.3.1. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

> *Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.*

Table 1.11.3.1: Strong Acids and Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
$HClO_3$	$Mg(OH)_2$
$HClO_4$	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table 1.11.3.1); any base not listed is a weak base. All strong bases are OH^- compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

Example 1.11.3.1: Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$
c. C_5H_5N

Solution

a. Because HCl is listed in Table 1.11.3.1, it is a strong acid.
b. Because $Mg(OH)_2$ is listed in Table 1.11.3.1, it is a strong base.
c. The nitrogen in C_5H_5N would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

Exercise 1.11.3.1

Identify each acid or base as strong or weak.

a. RbOH
b. HNO_2

Answer a

strong base

Answer b

weak acid

Example 1.11.3.2: Characterizing Base Ionization

Write the balanced chemical equation for the dissociation of $Ca(OH)_2$ and indicate whether it proceeds 100% to products or not.

Solution

This is an ionic compound of Ca^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:

$$\mathrm{Ca(OH)_2 \rightarrow Ca^{2+}(aq) + 2\,OH^-(aq)}$$

Because $Ca(OH)_2$ is listed in Table 1.11.3.1, this reaction proceeds 100% to products.

? Exercise 1.11.3.2

Write the balanced chemical equation for the dissociation of hydrazoic acid (HN_3) and indicate whether it proceeds 100% to products or not.

Answer a

The reaction is as follows:

$$HN_3 \rightarrow H^+(aq) + N_3^-(aq)$$

It does not proceed 100% to products because hydrazoic acid is not a strong acid.

Key Takeaways

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.
- Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions.

Contributions & Attributions

1.11.4: Strong and Weak Acids and Acid Ionization Constant (K a) (Ka)

The etching of glass is a slow process that can produce beautiful artwork. Traditionally, the glass has been treated with dilute hydrofluoric acid which gradually dissolves the glass under it. Parts of the piece that should not be etched are covered with wax or some other non-reactive material. In more recent times, compounds such as ammonium bifluoride have been used. Whichever chemical is employed, the artist must be very careful not to get any on their skin.

Strong and Weak Acids and Acid Ionization Constant

Acids are classified as either strong or weak, based on their ionization in water. A **strong acid** is an acid which is completely ionized in an aqueous solution. Hydrogen chloride (HCl) ionizes completely into hydrogen ions and chloride ions in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^+\,(aq) + \mathrm{Cl}^-\,(aq)$$

A **weak acid** is an acid that ionizes only slightly in an aqueous solution. Acetic acid (found in vinegar) is a very common weak acid. Its ionization is shown below.

$$\mathrm{CH_3COOH}\,(aq) \rightleftharpoons \mathrm{H}^+\,(aq) + \mathrm{CH_3COO}^-\,(aq)$$

The ionization of acetic acid is incomplete, and so the equation is shown with a double arrow. The extent of ionization of weak acids varies, but is generally less than 10%. A $0.10\ \mathrm{M}$ solution of acetic acid is only about 1.3% ionized, meaning that the equilibrium strongly favors the reactants.

Weak acids, like strong acids, ionize to yield the H^+ ion and a conjugate base. Because HCl is a strong acid, its conjugate base $\left(\mathrm{Cl}^-\right)$ is extremely weak. The chloride ion is incapable of accepting the H^+ ion and becoming HCl again. In general, the stronger the acid, the weaker its conjugate base. Likewise, the weaker the acid, the stronger its conjugate base.

Table 1.11.4.1: Relative Strengths of Acids and their Conjugate Bases

Acid	Conjugate Base
Strong Acids	
HCl (hydrochloric acid) (strongest)	Cl^- (chloride ion) (weakest)
$\mathrm{H_2SO_4}$ (sulfuric acid)	$\mathrm{HSO_4^-}$ (hydrogen sulfate ion)
$\mathrm{HNO_3}$ (nitric acid)	$\mathrm{NO_3^-}$ (nitrate ion)
Weak Acids	
$\mathrm{H_3PO_4}$ (phosphoric acid)	$\mathrm{H_2PO_4^-}$ (dihydrogen phosphate ion)
$\mathrm{CH_3COOH}$ (acetic acid)	$\mathrm{CH_3COO}^-$ (acetate ion)
$\mathrm{H_2CO_3}$ (carbonic acid)	$\mathrm{HCO_3^-}$ (hydrogen carbonate ion)
HCN (hydrocyanic acid) (weakest)	CN^- (cyanide ion) (strongest)

Strong acids are 100% ionized in solution. Weak acids are only slightly ionized. Phosphoric acid is stronger than acetic acid, and so is ionized to a greater extent. Acetic acid is stronger than carbonic acid, and so on.

The Acid Ionization Constant, K_a

The ionization for a general weak acid, HA, can be written as follows:

$$\mathrm{HA}\,(aq) \rightarrow \mathrm{H}^+\,(aq) + \mathrm{A}^-\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_\mathrm{a} = \frac{\left[\mathrm{H}^+\right]\left[\mathrm{A}^-\right]}{[\mathrm{HA}]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution. Therefore, the numerical value of K_a is a reflection of the strength of the acid. Weak acids with relatively higher K_a values are stronger than acids with relatively lower K_a values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the K_a value approaches infinity. For this reason, K_a values are generally reported for weak acids only.

The table below is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table 1.11.4.2: Acid Ionization Constants at 25°C

Name of Acid	Ionization Equation	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$ $HSO_4 \rightleftharpoons H^+ + SO_4^{2-}$	very large 1.3×10^{-2}
Oxalic acid	$H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$ $HC_2O_4 \rightleftharpoons H^+ + C_2O_4^{2-}$	6.5×10^{-2} 6.1×10^{-5}
Phosphoric acid	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ $H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2--}$ $HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$	7.5×10^{-3} 6.2×10^{-8} 4.8×10^{-13}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.2×10^{-7} 4.8×10^{-11}
Hydrocyanic acid	$HCN \rightleftharpoons H^+ + CN^-$	4.9×10^{-10}

Summary

- A strong acid is an acid which is completely ionized in an aqueous solution.
- A weak acid is an acid that ionizes only slightly in an aqueous solution.
- The acid ionization constant (K_a) is defined.

1.11.4.1: Practice Acids and Bases

1.11.5: Water - Acid and Base in One

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H_3O^+ and OH^- in aqueous solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4^+ + OH^-$$

where H_2O acts as an acid (in red).

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^-$$

where H_2O acts as an base (in blue).

It may not surprise you to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{1.11.5.1}$$

This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**.

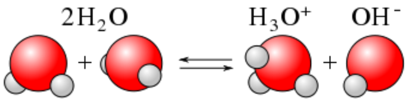


Figure 1.11.5.1: Autoionization of water, resulting in hydroxide and hydronium ions.

At this level, the concentration of both $H_3O^+(aq)$ and $OH^-(aq)$ in a sample of pure H_2O is about $1.0 \times 10^{-7}\ M$ (at room temperature). If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7} \tag{1.11.5.2}$$

for *any* sample of pure water because H_2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14}:

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

- For acids, the concentration of $H_3O^+(aq)$ (i.e., $[H_3O^+]$) is greater than $1.0 \times 10^{-7}\ M$.
- For bases the concentration of $OH^-(aq)$ (i.e., $[OH^-]$) is greater than $1.0 \times 10^{-7}\ M$.

However, the *product* of the two concentrations—$[H_3O^+][OH^-]$—is *always* equal to 1.0×10^{-14}, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{1.11.5.3}$$

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$) has to be for the product to equal 1.0×10^{-14}; or if you know $[OH^-]$), you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

Warning: Temperature Matters

The degree of autoionization of water (Equation 1.11.5.1)—and hence the value of K_w—changes with temperature, so Equations 1.11.5.2- 1.11.5.3are accurate only at room temperature.

Example 1.11.5.1: Hydroxide Concentration

What is $[\mathrm{OH^-}]$) of an aqueous solution if $[\mathrm{H_3O^+}]$ is $1.0 \times 10^{-4}\,M$?

Solution

Solutions to Example 14.7.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[\mathrm{H_3O^+}] = 1.0 \times 10^{-4}\ M$ Find: $[\mathrm{OH^-}]$ = ? M
List other known quantities.	none
Plan the problem.	Using the expression for K_w, (Equation 1.11.5.3), rearrange the equation algebraically to solve for $[\mathrm{OH^-}]$. $[\mathrm{OH^-}] = \dfrac{1.0 \times 10^{-14}}{[H_3O^+]}$
Calculate.	Now substitute the known quantities into the equation and solve. $[\mathrm{OH^-}] = \dfrac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10}\,M$ It is assumed that the concentration unit is molarity, so $[\mathrm{OH^-}]$ is 1.0 $\times 10^{-10}$ M.
Think about your result.	The concentration of the acid is high (> 1 x 10^{-7} M), so $[\mathrm{OH^-}]$ should be low.

Exercise 1.11.5.1

What is $[\mathrm{OH^-}]$ in a 0.00032 M solution of H_2SO_4?

Hint

Assume **both** protons ionize from the molecule...although this is not the case.

Answer

$3.1 \times 10^{-11}\ M$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H_3O^+ or OH^- ions in the formula unit because $[\mathrm{H_3O^+}]$ or $[\mathrm{OH^-}]$) may not be the same as the concentration of the acid or base itself.

Example 1.11.5.2: Hydronium Concentration

What is $[\mathrm{H_3O^+}]$ in a 0.0044 M solution of $\mathrm{Ca(OH)_2}$?

Solution

Solutions to Example 14.7.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[\mathrm{Ca(OH)_2}] = 0.0044\ M$ Find: $[\mathrm{H_3O^+}]$ = ? M

Steps for Problem Solving	
List other known quantities.	We begin by determining $[\mathrm{OH^-}]$. The concentration of the solute is 0.0044 M, but because $\mathrm{Ca(OH)_2}$ is a strong base, there are two $\mathrm{OH^-}$ ions in solution for every formula unit dissolved, so the actual $[\mathrm{OH^-}]$ is two times this: $[\mathrm{OH^-}]{=}2\ \times 0\ \cdot 0044\ \mathrm{M}{=}0\ \cdot 0088\ \mathrm{M}\cdot$
Plan the problem.	Use the expression for K_w (Equation 1.11.5.3) and rearrange the equation algebraically to solve for $[\mathrm{H_3O^+}]$. $[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{[OH^-]}$
Calculate.	Now substitute the known quantities into the equation and solve. $[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$ $[\mathrm{H_3O^+}]$ has decreased significantly in this basic solution.
Think about your result.	The concentration of the base is high (> 1 x 10^{-7} M) so $[\mathrm{H_3O^+}]$ should be low.

? Exercise 1.11.5.2

What is $[\mathrm{H_3O^+}]$ of an aqueous solution if $[\mathrm{OH^-}]$ is $1.0 \times 10^{-9}\ M$?

Answer

1.0×10^{-5} M

In any aqueous solution, the product of $[\mathrm{H_3O^+}]$ and $[\mathrm{OH^-}]$ equals 1.0×10^{-14} (at room temperature).

Contributions & Attributions

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1.11.6: The pH and pOH Scales - Ways to Express Acidity and Basicity

Learning Objectives

- Define pH and pOH.
- Determine the pH of acidic and basic solutions.
- Determine the hydronium ion concentration and pOH from pH.

As we have seen, $[H_3O^+]$ and $[OH^-]$ values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of $[H_3O^+]$:

$$pH = -\log[H_3O^+] \tag{1.11.6.1}$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H_3O^+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the pH scale. The pH scale is the range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. Figure $1.11.6.1$ illustrates this relationship, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H_3O^+]$, which will give a positive value for pH.

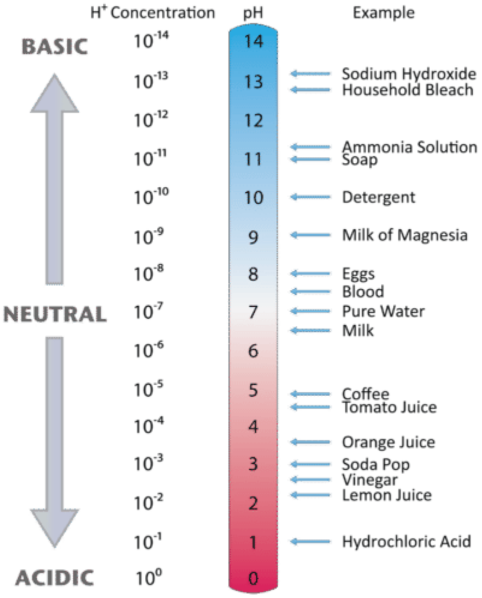


Figure 1.11.6.1: The pH values for several common materials.

Example 1.11.6.1

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Answer

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.

c. With a pH of less than 7, wine is acidic.

Exercise 1.11.6.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

a. human blood with pH = 7.4
b. household ammonia with pH = 11.0
c. cherries with pH = 3.6

Answer a

basic

Answer b

basic

Answer c

acidic

Calculating pH from Hydronium Concentration

The pH of solutions can be determined by using logarithms as illustrated in the next example for stomach acid. Stomach acid is a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}\ M$, what is the pH of the solution?

$$\begin{aligned} \text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) = 2.92 \end{aligned}$$

Logarithms

To get the **log value** on your calculator, enter the number (in this case, the hydronium ion concentration) first, then press the LOG key.

If the number is 1.0 x 10^{-5} (for $[H_3O^+]$ = 1.0 x 10^{-5} M) you should get an answer of "-5".

If you get a different answer, or an error, try pressing the LOG key before you enter the number.

Example 1.11.6.2: Converting Ph to Hydronium Concentration

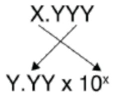
Find the pH, given the $[H_3O^+]$ of the following:

a. 1×10^{-3} M
b. 2.5×10^{-11} M
c. 4.7×10^{-9} M

Solution

Solutions to Example 14.9.2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: a. $[H_3O^+] = 1 \times 10^{-3}$ M b. $[H_3O^+] = 2.5 \times 10^{-11}$ M c. $[H_3O^+] = 4.7 \times 10^{-9}$ M Find: ? pH
Plan the problem.	Need to use the expression for pH (Equation 1.11.6.1). $pH = -\log [H_3O^+]$

Steps for Problem Solving	
Calculate.	Now substitute the known quantity into the equation and solve. a. $\text{pH} = -\log[\underline{1} \times 10^{-3}] = 3.\underline{0}$ (1 decimal places since 1 has 1 significant figure) b. $\text{pH} = -\log[\underline{2.5} \times 10^{-11}] = 10.\underline{60}$ (2 decimal places since 2.5 has 2 significant figures) c. $\text{pH} = -\log[\underline{4.7} \times 10^{-9}] = 8.\underline{30}$ (2 decimal places since 4.7 has 2 significant figures) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 1.11.6.2

Find the pH, given $[H_3O^+]$ of the following:

a. 5.8×10^{-4} M
b. 1.0×10^{-7}

Answer a

3.22

Answer b

7.00

Calculating Hydronium Concentration from pH

Sometimes you need to work "backwards"—you know the pH of a solution and need to find $[H_3O^+]$, or even the concentration of the acid solution. How do you do that? To convert pH into $[H_3O^+]$ we solve Equation 1.11.6.1 for $[H_3O^+]$. This involves taking the antilog (or inverse log) of the negative value of pH .

$$[\mathrm{H_3O^+}] = \mathrm{antilog}(-pH)$$

or

$$[\mathrm{H_3O^+}] = 10^{-pH} \tag{1.11.6.2}$$

As mentioned above, different calculators work slightly differently—make sure you can do the following calculations using ***your*** calculator.

Calculator Skills

We have a solution with a pH = 8.3. What is $[H_3O^+]$?

With some calculators you will do things in the following order:

1. Enter 8.3 as a negative number (use the key with both the +/- signs, not the subtraction key).
2. Use your calculator's 2nd or Shift or INV (inverse) key to type in the symbol found above the LOG key. The shifted function should be 10^x.
3. You should get the answer $\mathbf{5.0 \times 10^{-9}}$.

Other calculators require you to enter keys in the order they appear in the equation.

1. Use the Shift or second function to key in the 10^x function.

2. Use the +/- key to type in a negative number, then type in 8.3.
3. You should get the answer **5.0 × 10^{-9}**.

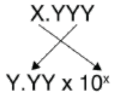
If neither of these methods work, try rearranging the order in which you type in the keys. Don't give up—you must master your calculator!

Example 1.11.6.3: Calculating Hydronium Concentration from pH

Find the hydronium ion concentration in a solution with a pH of 12.6. Is this solution an acid or a base? How do you know?

Solution

Solutions to Example 14.9.3

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: pH = 12.6 Find: $[H_3O^+]$ = ? M
Plan the problem.	Need to use the expression for $[H_3O^+]$ (Equation 1.11.6.2). $[H_3O^+]$ = antilog (-pH) or $[H_3O^+] = 10^{-pH}$
Calculate.	Now substitute the known quantity into the equation and solve. $[H_3O^+]$ = antilog ($12.\underline{60}$) = $\underline{2.5}$ x 10^{-13} M (2 significant figures since 4.7 has 12.60 2 decimal places) or $[H_3O^+] = 10^{-12.\underline{60}}$ = $\underline{2.5}$ x 10^{-13} M (2 significant figures since 4.7 has 12.60 2 decimal places) The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer: 

Exercise 1.11.6.3

If moist soil has a pH of 7.84, what is $[H_3O^+]$ of the soil solution?

Answer

1.5 x 10^{-8} M

The pOH scale

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration.

$$\text{pOH} = -\log\left[\text{OH}^-\right]$$

The relation between the hydronium and hydroxide ion concentrations expressed as p-functions is easily derived from the K_w expression:

$$K_\text{w} = [\text{H}_3\text{O}^+][\text{OH}^-] \tag{1.11.6.3}$$

$$-\log K_\text{w} = -\log([\text{H}_3\text{O}^+][\text{OH}^-]) = -\log[\text{H}_3\text{O}^+] + -\log[\text{OH}^-] \tag{1.11.6.4}$$

$$\text{p}K_\text{w} = \text{pH} + \text{pOH} \tag{1.11.6.5}$$

At 25 °C, the value of K_w is 1.0×10^{-14}, and so:

$$14.00 = \text{pH} + \text{pOH} \tag{1.11.6.6}$$

The hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7}\ M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{1.11.6.7}$$

$$\mathrm{pOH} = -\log[\mathrm{OH^-}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{1.11.6.8}$$

And so, at this temperature, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities less than $1.0 \times 10^{-7}\ M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7}\ M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

✓ Example 1.11.6.4:

Find the pOH of a solution with a pH of 4.42.

Solution

Solutions to Example 14.9.4

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: pH =4.42 Find: ? pOH
Plan the problem.	Need to use the expression pOH = 14 - pH
Calculate.	Now substitute the known quantity into the equation and solve. pOH=14−4.42=9.58
Think about your result.	The pH is that of an acidic solution, and the resulting pOH is the difference after subtracting from 14. The answer has two significant figures because the given pH has two decimal places.

? Exercise 1.11.6.4

The pH of a solution is 8.22. What is the pOH?

Answer

5.78

The diagram below shows all of the interrelationships between [H3O+][H3O+], [OH−][OH−], pH, and pOH.

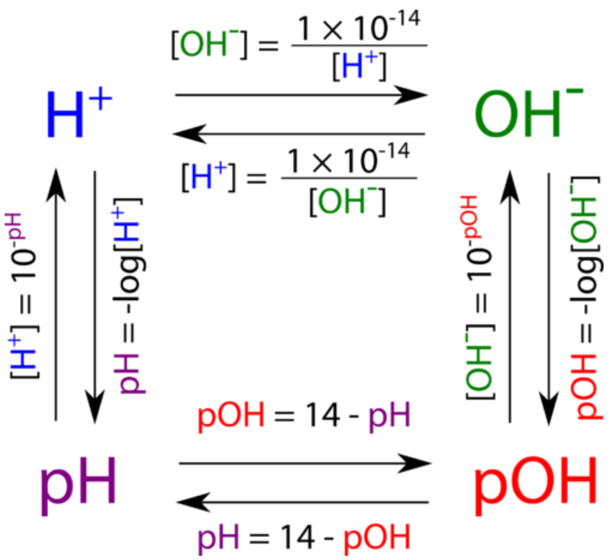


Figure 1.11.6.1: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH.

Contributions & Attributions

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-
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1.11.6.1: Practice pH

1.11.7: Reactions of Acids and Bases

Learning Objectives

- Write acid-base neutralization reactions.
- Write reactions of acids with metals.
- Write reactions of bases with metals.

Neutralization Reactions

The reaction that happens when an acid, such as HCl, is mixed with a base, such as $NaOH$:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than H^+ and a negative ion other than the hydroxide ion, OH^-. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

After removing the spectator ions, we get the net ionic equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

When a strong acid and a strong base are combined in the proper amounts—when $[H^+]$ equals $[OH^-$]\)—a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic.

Video: Equimolar (~0.01 M) and equivolume solutions of HCl and $NaOH$ are combined to make salt water. https://youtu.be/TS-I9KrUjB0

Example 1.11.7.1: Propionic Acid + Calcium Hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$].

Solution

Solutions to Example 14.5.1

Steps	Reaction

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$
Balance the equation. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$, to provide H^+ ions.	$\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$

? Exercise 1.11.7.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer

$$Ba(OH)_2(s) + 2\,CH_3CO_2H(aq) \rightarrow Ba(CH_3CO_2)_2(aq) + 2\,H_2O(l)$$

Acids and Bases React with Metals

Acids react with most metals to form a salt and hydrogen gas. As discussed previously, metals that are more active than acids can undergo a **single displacement reaction**. For example, zinc metal reacts with hydrochloric acid, producing zinc chloride and hydrogen gas.

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Bases also react with certain metals, like zinc or aluminum, to produce hydrogen gas. For example, sodium hydroxide reacts with zinc and water to form sodium zincate and hydrogen gas.

$$Zn(s) + 2\,NaOH(aq) + 2\,H_2O(l) \rightarrow Na_2Zn(OH)_4(aq) + H_2(g).$$

1.11.8: Acid–Base Titration

Note from Dr. B.

We will be looking at acids and bases more in Module 7, but a titration is a quantitative way of looking at solution concentrations of acids and bases, so we can look at it in Module 6.

Don't worry about the very end of the page (Indicator Selection for Titrations) because we haven't covered pH yet.

Learning Objectives

- Understand the basics of acid-base titrations.
- Understand the use of indicators.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base. The indicator will signal, by color change, when the base has been neutralized (when $[H^+] = [OH^-]$). At that point—called the **equivalence point,** or **end point**—the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

It is important to accurately measure volumes when doing titrations. The instrument you would use is called a burette (or buret).

Figure 1.11.8.1: Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\#\text{ mol HCl} = (0.02566\text{ L})(0.1078\text{ M}) = 0.002766\text{ mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:

$$\mathrm{HCl + NaOH \rightarrow NaCl + H_2O}$$

So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766\ mol\ \cancel{HCl} \times \frac{1\ mol\ NaOH}{1\ mol\ \cancel{HCl}} = 0.002766\ mol\ NaOH$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766\ mol\ \cancel{HCl} \times \frac{40.00\ g\ NaOH}{1\ mol\ \cancel{HCl}} = 0.1106\ g\ NaOH$$

This type of calculation is performed as part of a titration.

Example 1.11.8.1: Equivalence Point

What mass of $Ca(OH)_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3? The balanced chemical equation is as follows:

$$2\,HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2\,H_2O$$

Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

moles HNO_3 = (0.04402 L)(0.0885 M) = 0.00390 mol HNO_3

Using the balanced chemical equation, we can determine the number of moles of $Ca(OH)_2$ present in the analyte:

$$0.00390\ mol\ \cancel{HNO_3} \times \frac{1\ mol\ Ca(OH)_2}{2\ mol\ \cancel{HNO_3}} = 0.00195\ mol\ Ca(OH)_2$$

Then we convert this to a mass using the molar mass of $Ca(OH)_2$:

$$0.00195\ mol\ \cancel{Ca(OH)_2} \times \frac{74.1\ g\ Ca(OH)_2}{mol\ \cancel{Ca(OH)_2}} = 0.144\ g\ Ca(OH)_2$$

Exercise 1.11.8.1

What mass of $H_2C_2O_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:

$$H_2C_2O_4 + 2\,NaOH \rightarrow Na_2C_2O_4 + 2\,H_2O$$

Answer

0.182 g

Exercise 1.11.8.2

If 25.00 mL of HCl solution with a concentration of 0.1234 M is neutralized by 23.45 mL of NaOH, what is the concentration of the base?

Answer

0.1316 M NaOH

Exercise 1.11.8.3

A 20.0 mL solution of strontium hydroxide, $Sr(OH)_2$, is placed in a flask and a drop of indicator is added. The solution turns color after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the $Sr(OH)_2$ solution?

Answer

$3.12 \times 10^{-2}\ M$ $Sr(OH)_2$

Indicator Selection for Titrations

The indicator used depends on the type of titration performed. The indicator of choice should change color when enough of one substance (acid or base) has been added to exactly use up the other substance. Only when a strong acid and a strong base are produced will the resulting solution be neutral. The three main types of acid-base titrations, and suggested indicators, are:

The three main types of acid-base titrations, suggested indicators, and explanations

Titration between . . .	Indicator	Explanation
strong acid and strong base	any	
strong acid and weak base	methyl orange	changes color in the acidic range (3.2 - 4.4)
weak acid and strong base	phenolphthalein	changes color in the basic range (8.2 - 10.6)

Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all the analyte has reacted with the titrant.

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1.11.8.1: Practice Titration Calculations

1.11.9: Buffers are Solutions that Resist pH Change

Learning Objective

- Define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to life. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

This mechanism involves a buffer, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia (NH_3, a weak base) and ammonium chloride (NH_4Cl, a salt derived from that base).

Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base—a source of $OH^-(aq)$ ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(\ell) + C_2H_3O_2^-(aq) \tag{1.11.9.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

> Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq) \tag{1.11.9.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure 1.11.9.1illustrates both actions of a buffer.

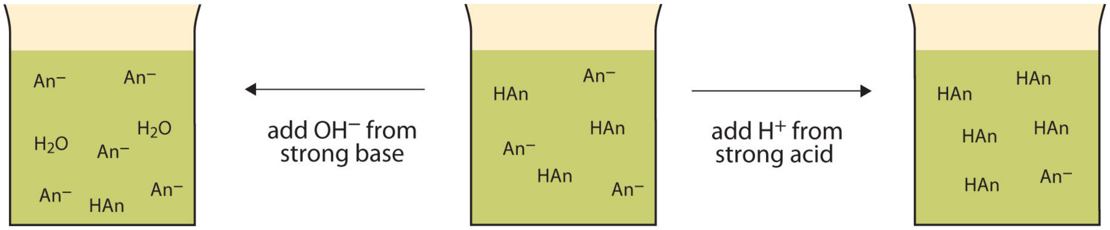

Figure 1.11.9.1: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq) \tag{1.11.9.3}$$

while the **ammonium ion** ($NH_4^+(aq)$) can react with any hydroxide ions introduced by strong bases:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(\ell) \tag{1.11.9.4}$$

✓ Example 1.11.9.1: Making Buffer Solutions

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $HCHO_2$ and $NaCHO_2$
b. HCl and NaCl
c. CH_3NH_2 and CH_3NH_3Cl
d. NH_3 and NaOH

Solution

a. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid—the formate ion (CHO_2^-). The combination of these two solutes would make a buffer solution.
b. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
c. Methylamine (CH_3NH_2) is like ammonia with one of its hydrogen atoms substituted with a CH_3 (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound CH_3NH_3Cl is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
d. Ammonia (NH_3) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

? Exercise 1.11.9.1

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. H_3PO_4 and NaH_2PO_4
c. NH_3 and $(NH_4)_3PO_4$
d. NaOH and NaCl

Answer a

Yes.

Answer b

No. Need a weak acid or base and a salt of its conjugate base or acid.

Answer c

Yes.

Answer d

No. Need a weak base or acid.

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

> A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/− type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.
>
> Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

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1.11.9.1: Practice Buffer Identification

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