CH104: Allied Health Chemisty I (2nd Edition)

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1: Introduction to Chemistry and the Scientific Method

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1.1: What is Chemistry?

> **◗ Learning Objectives**
>
> - Understand of the scope of chemistry as a scientific discipline.

Matter is any substance that has mass and takes up space. Matter includes atoms and anything made up of atoms. It does not include various forms of energy such as light, sound, or heat. Chemists often view matter on two characteristic length scales: the macroscopic and the nanoscopic. The nanoscopic level can also be called the particle level.

Figure 1.1.1: Sun shinning through the trees in a forest.

The macroscopic scale is the length scale on which objects or phenomena are large enough to sense directly. Everything that one can see, touch, and handle in the forest of Figure 1.1.1 is within the macroscopic scale. However, each of these items can be decomposed into smaller nanoscopic scale particles.

The nanoscopic scale is the scale of atoms and molecules. All of the everyday objects that we can see or touch are ultimately composed of atoms. This ordinary atomic matter is in turn made up of interacting subatomic particles—usually a nucleus of protons and neutrons, and a cloud of orbiting electrons. Chemistry is an attempt to understand how the nanoscopic level of matter controls the macroscopic level of matter.

> **📌 Matter vs. Mass**
>
> Matter should not be confused with mass, as the two are not the same in modern physics. Matter is a physical substance of which systems may be composed, while mass is not a substance, but rather a quantitative property of matter and other substances or systems.

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

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). 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.
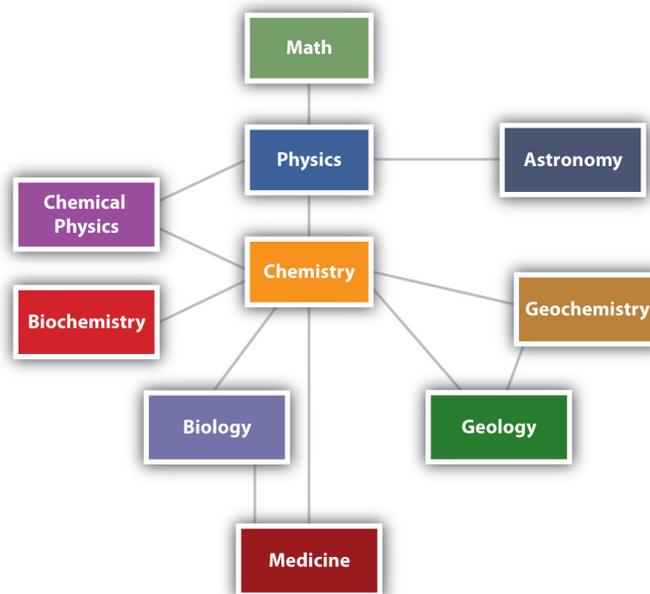
Figure 1.1.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

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 shows how many of the individual fields of science are related.

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1.2: The Scientific Method

> **◑ Learning Objectives**
>
> - Define the steps of the scientific method.

How do scientists work? Generally, they follow a process called the scientific method. The scientific method is an organized procedure for learning answers to questions and making explanations for observations. To find the answer to a question (for example, "Why do birds fly toward Earth's equator during the cold months?"), a scientist goes through the following steps, which are also illustrated in Figure 1.2.2:
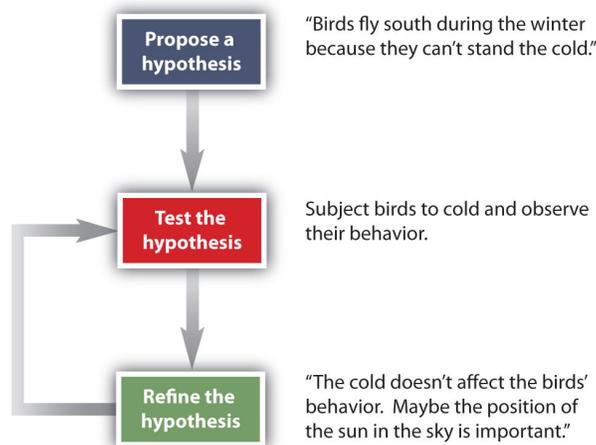
Figure 1.2.2: The General Steps of the Scientific Method. After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.

The steps may not be as clear-cut in real life as described here, but most scientific work follows this general outline.

1. **Propose a hypothesis.** A scientist generates a testable idea, or hypothesis, to try to answer a question or explain an observation about how the natural universe works. Some people use the word *theory* in place of hypothesis, but the word hypothesis is the proper word in science. For scientific applications, the word theory is a general statement that describes a large set of observations and data. A theory represents the highest level of scientific understanding.
2. **Test the hypothesis.** A scientist evaluates the hypothesis by devising and carrying out experiments to test it. If the hypothesis passes the test, it may be a proper answer to the question. If the hypothesis does not pass the test, it may not be a good answer.
3. **Refine the hypothesis if necessary.** Depending on the results of experiments, a scientist may want to modify the hypothesis and then test it again. Sometimes the results show the original hypothesis to be completely wrong, in which case a scientist will have to devise a new hypothesis.

Not all scientific investigations are simple enough to be separated into these three discrete steps. But these steps represent the general method by which scientists learn about our natural universe.

1.3: Hypothesis, Theories, and Laws

Learning Objectives

- Define the scientific means of hypothesis, theory, and 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.

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1.4: Types of Observations

> ◐ **Learning Objectives**
>
> - Distinguish between quantitative and qualitative 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 concentrated 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 °C, and 35.9 grams of table salt—whose chemical name is sodium chloride—dissolve in 100 grams of water at 20 °C.

> **? Exercise 1.4.1**
>
> Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.
>
> a. 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.
> b. Litmus paper dipped in lemon juice turns red.
> c. A prism separates white light into a spectrum of colors.
> d. Limestone is relatively insoluble in water but dissolves readily in dilute acid with the evolution of a gas.
> e. Gas mixtures that contain more than 4% hydrogen in air are potentially explosive.
>
> **Answer a**
>
> quantitative observation
>
> **Answer b**
>
> qualitative observation
>
> **Answer c**
>
> qualitative observation
>
> **Answer d**
>
> qualitative observation
>
> **Answer e**
>
> quantitative observation

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1.5: Experimental Design

> **◑ Learning Objectives**
>
> - Classify dependent, independent, and controlled variables.

An **experiment** is a controlled method of testing a hypothesis. Experiments are designed to provide an opportunity to make observation that will help test a hypothesis. Experiments are best understood in term of three types of variables: independent variables, dependent variables, and controlled variables. It may be helpful to think of a variable simply as something that can be measured.

Most experiments are repeated multiple times with slight variations. These repetitions are often called trials. A variable that is purposely altered between trials is called an **independent variable**. Usually, it is a best practice to have a single independent variable in an experiment. It may be helpful to think of the independent variable as the "input" of the experiment.

If the independent variable is the "input" of an experiment, than the **dependent variable** is the "output." Dependent variables change in response to the independent variable (their name comes from the fact that they depend on the independent variable).

Finally, some variables, called **controlled variables**, are kept constant through all of the trials. Controlled variables are kept constant so that their fluctuations do not alter the dependent variable and cloud its relationship with the independent variable.

Sometimes a scatter plot (a mathematical diagram in which data is graphed against an x and a y axis) is used to display results from an experiment. If a scatter plot is used, the independent variable is generally plotted on the horizontal axis and the dependent variable is plotted on the vertical axis.

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2: Measurement and Significant Figures

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2.1: Measurements

> **Learning Objectives**
>
> - Express quantities properly using a number and a unit.

A coffee maker's instructions tell you to fill the coffeepot with 4 cups of water and use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure 2.1.1); the nurse is also measuring.

Figure 2.1.1: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GNU Free Documentation License; Pia von Lützau via Wikipedia).

Chemists measure the properties of matter using a variety of devices or measuring tools, many of which are similar to those used in everyday life. Rulers are used to measure length, balances (scales) are used to measure mass (weight), and graduated cylinders or pipettes are used to measure volume. Measurements made using these devices are expressed as quantities. A **quantity** is an amount of something and consists of a **number** and a **unit**. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5.2 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5.2.

$$\underbrace{5.2}_{\text{number}}\ \underbrace{\text{kilometers}}_{\text{unit}}$$

If you ask a friend how far he or she walks from home to school, and the friend answers "12" without specifying a unit, you do not know whether your friend walks—for example, 12 miles, 12 kilometers, 12 furlongs, or 12 yards.

Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of "100" without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

> *Both a number and a unit must be included to express a quantity properly.*

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. The next two sections examine the rules for expressing numbers.

> **✔ Exercise 2.1.1**
>
> Identify the **number** and the **unit** in each quantity.
>
> a. one dozen eggs
> b. 2.54 centimeters
> c. a box of pencils
> d. 88 meters per second

Answer a

The number is one, and the unit is dozen.

Answer b

The number is 2.54, and the unit is centimeter.

Answer c

The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.

Answer d

The number is 88, and the unit is meters per second. Note that in this case the unit is actually a combination of two units: meters and seconds.

? Exercise 2.1.2

Identify the ***number*** and the ***unit*** in each quantity.

a. 99 bottles of soda
b. 60 miles per hour
c. 32 fluid ounces
d. 98.6 degrees Fahrenheit

Answer a

The number is 99, and the unit is bottles of soda.

Answer b

The number is 60, and the unit is miles per hour.

Answer c

The number 32, and the unit is fluid ounces

Answer d

The number is 98.6, and the unit is degrees Fahrenheit

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2.2: Scientific Notation

> ◉ **Learning Objectives**
>
> - Convert numbers between scientific notation and decimal form.

The instructions for making a pot of coffee specified 3 scoops (rather than 12,000 grounds) because any measurement is expressed more efficiently with units that are appropriate in size. In science, however, we often must deal with quantities that are extremely small or incredibly large. For example, you may have 5,000,000,000,000 red blood cells in a liter of blood, and the diameter of an iron atom is 0.000000014 inches. Numbers with many zeros can be cumbersome to work with, so scientists use scientific notation.

Scientific notation is a system for expressing very large or very small numbers in a compact manner. It uses the idea that such numbers can be rewritten as a simple number multiplied by 10 raised to a certain exponent, or power.

Let us look first at very large numbers. Suppose a spacecraft is 1,500,000 miles from Mars. The number 1,500,000 can be thought of as follows:

$$1.5 \times \underbrace{1,000,000}_{\underbrace{10 \times 10 \times 10 \times 10 \times 10 \times 10}_{10^6}} = 1.5 \times 10^6$$

That is, 1,500,000 is the same as 1.5 times 1 million, and 1 million is $10 \times 10 \times 10 \times 10 \times 10 \times 10$, or 10^6 (which is read as "ten to the sixth power"). Therefore, 1,500,000 can be rewritten as 1.5 times 10^6, or 1.5×10^6. The distance of the spacecraft from Mars can therefore be expressed as 1.5×10^6 miles.

Recall that:

- $10^0 = 1$
- $10^1 = 10$
- $10^2 = 100$
- $10^3 = 1,000$
- $10^4 = 10,000$
- and so forth

The standard convention for expressing numbers in scientific notation is to write a single *nonzero* first digit, a decimal point, and the rest of the digits, excluding any trailing zeros (see rules for significant figures in the next section for more details on what to exclude). This number is followed by a multiplication sign and then by 10 raised to the power necessary to reproduce the original number. For example, although 1,500,000 can also be written as 15. $\times 10^5$ (which would be 15. $\times 100,000$), the convention is to have only one digit before the decimal point. How do we know to what power 10 is raised? The power is the number of places you have to move the decimal point to the *left* to place it after the <u>first digit</u>, so that the number being multiplied is *between 1 and 10*:

$$1,500,000 = 1.5 \times 10^6$$

> ✔ **Example 2.2.1: Scientific Notation**
>
> Express each number in scientific notation.
>
> a. 67,000,000,000
> b. 1,689
> c. 12.6
>
> **Answer a**
>
> Moving the decimal point 10 places to the left gives 6.7×10^{10}.
>
> **Answer b**
>
> The decimal point is assumed to be at the end of the number, so moving it three places to the left gives 1.689×10^3.
>
> **Answer c**

In this case, we need to move the decimal point only one place to the left, which yields 1.26×10^1.

? Exercise 2.2.1

Express each number in scientific notation.

  a. 1,492
  b. 102,000,000
  c. 101,325

Answer a

  Moving the decimal point 3 places to the left gives 1.492×10^3.

Answer b

  The decimal point is assumed to be at the end of the number, so moving it 8 places to the left gives 1.02×10^8.

Answer c

  Moving the decimal point 5 places to the left yields 1.01325×10^5.

To change a number in **scientific notation** to **standard form**, we reverse the process, moving the decimal point to the right. Add zeros to the end of the number being converted, if necessary, to produce a number of the proper magnitude. Lastly, we drop the number 10 and its power.

$$1.5 \times 10^6 = 1.500000 = 1,500,000$$

✔ Example 2.2.2

Express each number in standard, or conventional notation.

  a. 5.27×10^4
  b. 1.0008×10^6

Answer a

  Moving the decimal four places to the right and adding zeros give 52,700.

Answer b

  Moving the decimal six places to the right and adding zeros give 1,000,800.

✔ Exercise 2.2.2

Express each number in standard, or conventional notation.

  a. 6.98×10^8
  b. 1.005×10^2

Answer a

  Moving the decimal point eight places to the right and adding zeros give 698,000,000.

Answer b

  Moving the decimal point two places to the right gives 100.5

We can also use scientific notation to express numbers whose magnitudes are less than 1. For example, the quantity 0.006 centimeters can be expressed as follows:

$$6 \times \underbrace{\frac{1}{1,000}} = 6 \times 10^{-3}$$

$$\underbrace{\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}}_{10^{-3}}$$

That is, 0.006 centimeters is the same as 6 *divided by* one thousand, which is the same as 6 *divided* by 10 x 10 x 10 or 6 *times* 10^{-3} (which is read as "ten to the negative third power"). Therefore, 0.006 centimeters can be rewritten as 6 times 10^{-3}, or 6×10^{-3} centimeters.

Recall that:

- $10^{-1} = 1/10$
- $10^{-2} = 1/100$
- $10^{-3} = 1/1,000$
- $10^{-4} = 1/10,000$
- $10^{-5} = 1/100,000$
- and so forth

We use a negative number as the power to indicate the number of places we have to move the decimal point to the right to make it follow the first nonzero digit so that the number is between 1 and 10. This is illustrated as follows:

$$0.006 = 6 \times 10^{-3}$$

> 📌 **Note:**
>
> In writing scientific notations, the convention is to have only one digit before the decimal point.
>
> - Numbers that are greater than one have a positive power in scientific notation. If the decimal point is moved to the left n places, the power (n) of 10 is positive. $1,500,000 = 1.5 \times 10^{6}$
> - Numbers that are less than one have a negative power in scientific notation. If the decimal point is moved to the right n places, the power (n) of 10 is negative. $0.006 = 6 \times 10^{-3}$

> ✔ **Example 2.2.3**
>
> Express each number in scientific notation.
>
> a. 0.000006567
> b. −0.0004004
> c. 0.000000000000123
>
> **Answer a**
>
> Move the decimal point six places to the right to get 6.567×10^{-6}.
>
> **Answer b**
>
> Move the decimal point four places to the right to get -4.004×10^{-4}. The negative sign on the number itself does not affect how we apply the rules of scientific notation.
>
> **Answer c**
>
> Move the decimal point 13 places to the right to get 1.23×10^{-13}.

> ❓ **Exercise 2.2.3**
>
> Express each number in scientific notation.
>
> a. 0.000355
> b. 0.314159

  c. −0.051204

Answer a

  Moving the decimal point four places to the right gives 3.55×10^{-4}.

Answer b

  Moving the decimal point one place to the right gives 3.14159×10^{-1}.

Answer c

  Moving the decimal point one place to the right gives -5.1204×10^{-2}.

As with numbers with positive powers of 10, when changing from **scientific** notation to **standard or conventional** format, we reverse the process.

$$6 \times 10^{-3} = 006. = 0.006$$

> 📌 Note
>
> Changing a number in scientific notation to standard form:
>
> - If the scientific notation has a positive power, the standard number is greater than one. Example: $8 \times 10^{4} = 80{,}000$
> - If the scientific notation has a negative power, then the standard number is less than one. Example: $8 \times 10^{-2} = 0.08$

> ✔ Example 2.2.4
>
> Change the number in scientific notation to standard form.
>
>   a. 6.22×10^{-2}
>   b. 9.9×10^{-9}
>
> **Answer a**
>
>   0.0622
>
> **Answer b**
>
>   0.0000000099

> ❓ Exercise 2.2.4
>
> Change the number in scientific notation to standard form.
>
>   a. 9.98×10^{-5}
>   b. 5.109×10^{-8}
>
> **Answer a**
>
>   0.0000998
>
> **Answer b**
>
>   0.00000005109

Although calculators can show 8 to 10 digits in their display windows, that is not always enough when working with very large or very small numbers. For this reason, many calculators are designed to handle scientific notation. The method for entering scientific notation differs for each calculator model, so take the time to learn how to do it properly on your calculator, *asking your instructor for assistance if necessary*. If you do not learn to enter scientific notation into your calculator properly, you will not get the correct final answer when performing a calculation.

2.3: Significant Figures

> ◖◗ **Learning Objectives**
>
> - Given an instrument and an object, make a measurement including units with the appropriate level of uncertainty.
> - Given a physical quantity, identify the correct number of significant figures.

Scientists have established certain conventions for communicating the degree of **precision** of a measurement, which is dependent on the measuring device used. Imagine, for example, that you are using a meterstick to measure the width of a table. The centimeters (cm) marked on the meterstick, tell you how many *centimeters* wide the table is. Many metersticks also have markings for millimeters (mm), so we can measure the table to the nearest *millimeter*. Most metersticks do not have any smaller (or more precise) markings indicated, so you cannot report the measured width of the table any more precise than to the nearest millimeter. However, you can *estimate* one past the smallest marking, in this case the millimeter, to the next decimal place in the measurement (Figure 2.3.1).
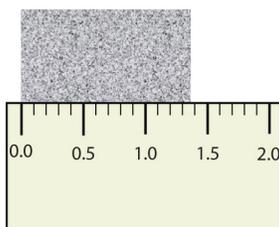
Figure 2.3.1: Measuring an Object to the Correct Number of Digits. How many digits should be reported for the length of this object?

The concept of **significant figures** takes this limitation into account. The significant figures of a measured quantity are defined as all the digits known with *certainty* (those indicated by the markings on the measuring device) ***and*** the first uncertain, or estimated, digit (one digit past the smallest marking on the measuring device). It makes no sense to report any digits after the first uncertain one, so it is the last digit reported in a measurement. Zeros are used when needed to place the significant figures in their correct positions. Thus, zeros are sometimes counted as significant figures but are sometimes only used as placeholders.

> | *"Sig figs" is a common abbreviation for significant figures.*

Consider the earlier example of measuring the width of a table with a meterstick. If the table is measured and reported as being 1,357 mm wide, the number 1,357 has four significant figures. The 1 (thousands place), the 3 (hundreds place), and the 5 (tens place) are certain; the 7 (ones place) is assumed to have been estimated. It would make no sense to report such a measurement as 1,357.0 (five Sig Figs) or 1,357.00 (six Sig Figs) because that would suggest the measuring device was able to determine the width to the nearest tenth or hundredth of a millimeter, when in fact it shows only tens of millimeters and therefore the ones place was estimated.

On the other hand, if a measurement is reported as 150 mm, the 1 (hundreds) and the 5 (tens) are known to be significant, but how do we know whether the zero is or is not significant? The measuring device could have had marks indicating every 100 mm or marks indicating every 10 mm. How can you determine if the zero is significant (the estimated digit), or if the 5 is significant and the zero a value placeholder?

The **rules** for deciding which digits in a measurement are significant are as follows:

1. All nonzero digits are significant. In 1,357 mm, all the digits are significant.
2. *Sandwiched (or embedded) zeros*, those between significant digits, are significant. Thus, 405 g has *three* significant figures.
3. *Leading zeros*, which are zeros at the beginning of a decimal number less than 1, are not significant. In 0.000458 mL, the first four digits are leading zeros and are not significant. The zeros serve only to put the digits 4, 5, and 8 in the correct decimal positions. This number has three significant figures.
4. *Trailing zeros*, which are zeros at the end of a number, are significant only if the number has a decimal point. Thus, in 1,500 m, the two trailing zeros are not significant because the number is written without a decimal point; the number has two significant figures. However, in 1,500.00 m, all six digits are significant because the number has a decimal point.

✔ **Example 2.3.1**

How many significant figures does each number have?

  a. 6,798,000
  b. 6,000,798
  c. 6,000,798.00
  d. 0.0006798

Answer a

  four (by rules 1 and 4)

Answer b

  seven (by rules 1 and 2)

Answer c

  nine (by rules 1, 2, and 4)

Answer d

  four (by rules 1 and 3)

? **Exercise 2.3.1**

How many significant figures does each number have?

  a. 2.1828
  b. 0.005505
  c. 55,050
  d. 5
  e. 500

Answer a

  five

Answer b

  four

Answer c

  four

Answer d

  one

Answer e

  one

Rounding off numbers

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 207.518m. 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 1.5.1.

Table 1.5.1: 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.

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2.4: Calculations and Significant Figures

> **◑ Learning Objectives**
>
> - When using measured physical quantities in calculations report the correct number of significant figures in the answer.

Combining Numbers

For **addition or subtraction**, the rule is to stack all the numbers with their decimal points aligned and then limit (round to) the answer's significant figures to the rightmost column for which all the numbers have significant figures. Consider the following:

$$
\begin{array}{r}
56.789 \\
+\quad 102.2 \\
+\ 1{,}300.099 \\
\hline
=\ 1{,}459.088
\end{array}
$$

↑ Limit to this column

The arrow points to the rightmost column in which all the numbers have significant figures—in this case, the tenths place. Therefore, we will limit our final answer to the tenths place. Is our final answer therefore 1,459.0? No, because when we drop digits from the end of a number, we also have to round the number. Notice that the first dropped digit, in the hundredths place, is 8. This suggests that the answer is actually closer to 1,459.1 than it is to 1,459.0, so we need to round up to **1,459.1**. The standard rules for rounding numbers are simple: If the first dropped digit is 5 or higher, round up. If the first dropped digit is lower than 5, do not round up.

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:

$$38.65 \times 105.93 = 4{,}094.1945$$

4 sig figs   5 sig figs   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 4,000 has one significant figure and should be written as the number 4×10^4. 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 reported digits are significant.

> **✔ Example 2.4.2**
>
> 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
> c. $1{,}027 + 610.0 + 363.06$
>
> **Answer a**
>
> 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 hundredths place) is greater than 5, we round up to 2,085.6, which in scientific notation is 2.0856×10^3.
>
> **Answer b**
>
> The calculator gives 1,125 as the answer, but we limit it to three significant figures and convert into scientific notation: 1.13×10^3.
>
> **Answer c**
>
> The calculator gives 2,000.06 as the answer, but because 1,027 has its farthest-right significant figure in the ones column, our answer must be limited to the ones position: 2,000 which in scientific notation is 2.000×10^3.

> **? Exercise 2.4.2**
>
> Write the answer for each expression using scientific notation with the appropriate number of significant figures.
>
> a. $217 \div 903$
> 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**
> $1437 = 1.437 \times 10^{3}$
> **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.

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2.5: The International System of Units

> **◐ Learning Objectives**
>
> - List SI units of length, mass, time, and temperature.
> - List the multipliers, symbols, and numerical meanings for all SI units.

People who live in the United States measure weight in pounds, height in feet and inches, and a car's speed in miles per hour. In contrast, chemistry and other branches of science use the International System of Units (also known as **SI** after *Système Internationale d'Unités*), which was established so that scientists around the world could communicate efficiently with each other. Many countries have also adopted SI units for everyday use as well. The United States is one of the few countries that has not.

Base SI Units

Base (or basic) units, are the fundamental units of SI. There are seven base units, which are listed in Table 2.5.1, Chemistry uses five of the base units: the mole for amount, the kilogram for mass, the meter for length, the second for time, and the Kelvin for temperature. The degree Celsius (°C) is also commonly used for temperature. The numerical relationship between Kelvins and degrees Celsius is as follows:

$$K = {}^\circ C + 273 \qquad\qquad (2.5.1)$$

Table 2.5.1: The Seven Base SI Units

Property	Unit	Abbreviation
length	meter	m
mass	kilogram	kg
time	second	s
amount	mole	mol
temperature	Kelvin	K
electrical current	ampere	amp
luminous intensity	candela	cd

> The United States uses the English (sometimes called Imperial) system of units for many quantities. Inches, feet, miles, gallons, pounds, and so forth, are all units connected with the English system of units. There have been many mistakes due to the improper conversion of units between the SI and English systems.

The size of each base unit is defined by international convention. For example, the *kilogram* is defined as the quantity of mass of a special metal cylinder kept in a vault in France (Figure 2.5.1). The other base units have similar definitions and standards. The sizes of the base units are not always convenient for all measurements. For example, a meter is a rather large unit for describing the width of something as narrow as human hair. Instead of reporting the diameter of hair as 0.00012 m or as 1.2×10^{-4} m using scientific notation as discussed in section 1.4, SI also provides a series of **prefixes** that can be attached to the units, creating units that are larger or smaller by powers of 10.

Figure 2.5.1: The Kilogram. The standard for the kilogram is a platinum-iridium cylinder kept in a special vault in France. Source: Photo reproduced by permission of the Bureau International des Poids et Mesures, who retain full internationally protected copyright.

Common prefixes and their multiplicative factors are listed in Table 2.5.2. (Perhaps you have already noticed that the base unit *kilogram* is a combination of a prefix, kilo- meaning 1,000 ×, and a unit of mass, the gram.) Some prefixes create a multiple of the original unit: 1 kilogram equals 1,000 grams, and 1 megameter equals 1,000,000 meters. Other prefixes create a fraction of the original unit. Thus, 1 centimeter equals 1/100 of a meter, 1 millimeter equals 1/1,000 of a meter, 1 microgram equals 1/1,000,000 of a gram, and so forth.

Table 2.5.2: Prefixes Used with SI Units

Prefix	Abbreviation	Multiplicative Factor	Multiplicative Factor in Scientific Notation
giga-	G	1,000,000,000 ×	$10^9 \times$
mega-	M	1,000,000 ×	$10^6 \times$
kilo-	k	1,000 ×	$10^3 \times$
deca-	D	10 ×	$10^1 \times$
deci-	d	1/10 ×	$10^{-1} \times$
centi-	c	1/100 ×	$10^{-2} \times$
milli-	m	1/1,000 ×	$10^{-3} \times$
micro-	μ*	1/1,000,000 ×	$10^{-6} \times$
nano-	n	1/1,000,000,000 ×	$10^{-9} \times$

*The letter μ is the Greek lowercase letter for *m* and is called "mu," which is pronounced "myoo."

Both SI units and prefixes have abbreviations, and the combination of a prefix abbreviation with a base unit abbreviation gives the abbreviation for the modified unit. For example, kg is the abbreviation for kilogram. We will be using these abbreviations throughout this book.

> 📌 The Difference Between Mass and Weight
>
> The mass of a body is a measure of its inertial property or how much matter it contains. The weight of a body is a measure of the force exerted on it by gravity or the force needed to support it. Gravity on earth gives a body a downward acceleration of about 9.8 m/s². In common parlance, weight is often used as a synonym for mass in weights and measures. For instance, the verb "to weigh" means "to determine the mass of" or "to have a mass of." The incorrect use of weight in place of mass should be phased out, and the term mass used when mass is meant. The SI unit of mass is the kilogram (kg). In science and technology, the weight of a body in a particular reference frame is defined as the force that gives the body an acceleration equal to the local acceleration of free fall in that reference frame. Thus, the SI unit of the quantity weight defined in this way (force) is the newton (N).

Derived SI Units

Derived units are combinations of SI base units. Units can be multiplied and divided, just as numbers can be multiplied and divided. For example, the area of a square having a side of 2 cm is 2 cm × 2 cm, or 4 cm^2 (read as "four centimeters squared" or "four square centimeters"). Notice that we have squared a length unit, the centimeter, to get a derived unit for area, the square centimeter.

Volume is an important quantity that uses a derived unit. **Volume** is the amount of space that a given substance occupies and is defined geometrically as length × width × height. Each distance can be expressed using the meter unit, so volume has the derived unit m × m × m, or m^3 (read as "meters cubed" or "cubic meters"). A cubic meter is a rather large volume, so scientists typically express volumes in terms of 1/1,000 of a cubic meter. This unit has its own name—the liter (L). A liter is a little larger than 1 US quart in volume. Below are approximate equivalents for some of the units used in chemistry.

Approximate Equivalents to Some SI Units

- 1 m ≈ 39.36 in. ≈ 3.28 ft ≈ 1.09 yd
- 1 in. = 2.54 cm (this is exact rather than an approximation)
- 1 km ≈ 0.6214 mi
- 1 kg ≈ 2.205 lb
- 1 lb ≈ 454 g
- 1 L ≈ 1.06 qt
- 1 qt ≈ 0.946 L
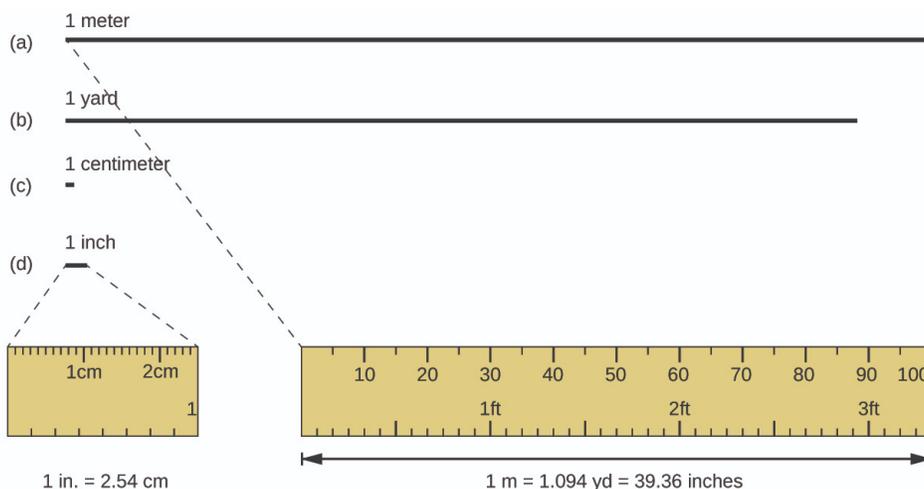

Figure 2.5.2: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd. (CC BY 4.0; OpenStax)

As shown in Figure 2.5.3, a liter is also 1,000 cm^3. By definition, there are 1,000 mL in 1 L, so 1 milliliter and 1 cubic centimeter represent the same volume.
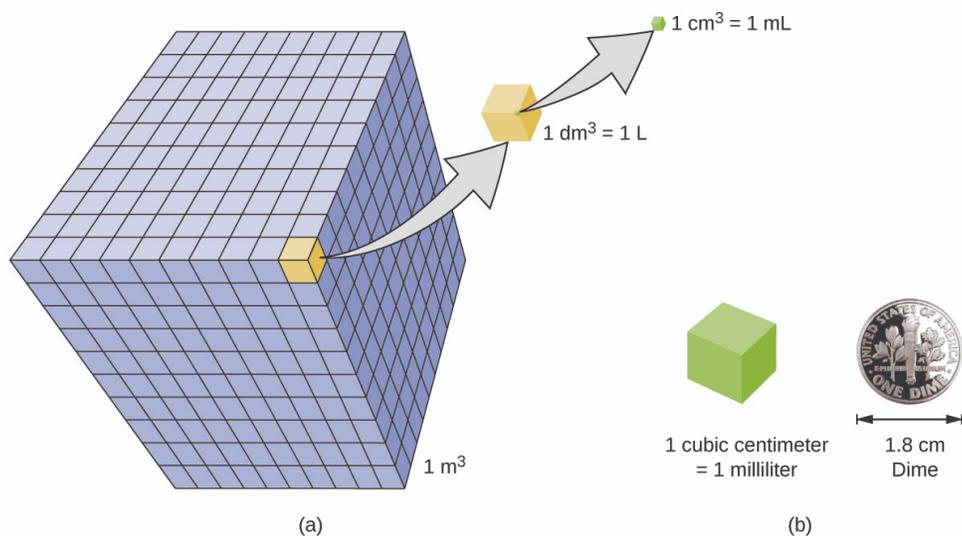
$$1 \; mL = 1 \; cm^3 \tag{2.5.2}$$

Figure $2.5.3$: Units of Volume. (a) The relative volumes are shown for cubes of 1 m³, 1 dm³ (1 L), and 1 cm³ (1 mL) (not to scale). A liter (L) is defined as a cube 1 dm (1/10th of a meter) on a side. A milliliter (mL), 1/1,000th of a liter, is equal to 1 cubic centimeter. (b) The diameter of a dime is compared relative to the edge length of a 1-cm³ (1-mL) cube. (CC BY 4.0; OpenStax)

✔ **Example 2.5.1**

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kiloliter
b. microsecond
c. decimeter
d. nanogram

Answer a

The abbreviation for a kiloliter is kL. Because kilo means "1,000 ×," 1 kL equals 1,000 L.

Answer b

The abbreviation for microsecond is μs. Micro implies 1/1,000,000th of a unit, so 1 μs equals 0.000001 s.

Answer c

The abbreviation for decimeter is dm. Deci means 1/10th, so 1 dm equals 0.1 m.

Answer d

The abbreviation for nanogram is ng and equals 0.000000001 g.

? **Exercise 2.5.1**

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kilometer
b. milligram
c. nanosecond
d. centiliter

Answer a

km (1,000 m)

Answer b

mg (0.001 g)

Answer c

ns (0.000000001 s)
Answer d
cL (0.01L)

Energy, another important quantity in chemistry, is the ability to perform work, such as moving a box of books from one side of a room to the other side. It has a derived unit of kg•m^2/s^2. (The dot between the kg and m units implies the units are multiplied together.) Because this combination is cumbersome, this collection of units is redefined as a **joule** (J). An older unit of energy, but likely more familiar to you, the calorie (cal), is also widely used. There are 4.184 J in 1 cal. Energy changes occur during all chemical processes and will be discussed in a later chapter.

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CHAPTER OVERVIEW

3: Dimensional Anlaysis and Density

3.1: Problem Solving and Unit Conversions

> **⊕ Learning Objectives**
>
> - Given a quantity, convert from one set of units to another using dimensional analysis showing canceling on units. This includes one factor conversions.

During your studies of chemistry (and physics also), you will note that mathematical equations are used in many different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Using the correct units is critical!

Converting Between Units with Conversion Factors

A **conversion factor** is a ratio used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, for example. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with.

Table 3.1.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 gallon (gal)	3.785 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table 3.1.1. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions in the SI system, as covered in the last chapter.
- Conversions in the English system (such as 12 inches = 1 foot or 4 quarts = 1 gallon).
- Time conversions (such as 60 seconds = 1 minute or 60 minutes = 1 hour).

If you learned the SI units and prefixes described, then you know that 1 cm is 1/100th of a meter.

$$1 \text{ cm} = \frac{1}{100} \text{ m} = 10^{-2} \text{m} \tag{3.1.1}$$

or

$$100 \text{ cm} = 1 \text{ m} \tag{3.1.2}$$

Suppose we divide both sides of the equation by 1m (both the number *and* the unit):

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1 \text{ m}}{1 \text{ m}} \tag{3.1.3}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1000 \text{ mm}}{1 \text{ m}} = \frac{1 \times 10^6 \,\mu\text{m}}{1 \text{ m}} = 1 \tag{3.1.4}$$

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units.

Performing Dimensional Analysis

Dimensional analysis is amongst the most valuable tools that physical scientists use. Simply put, it is the conversion between an amount in one unit to the corresponding amount in a desired unit using various conversion factors.

Here is a simple example. How many centimeters are there in 3.55 m? Perhaps you can determine the answer in your head. If there are 100 cm in every meter, then 3.55 m equals 355 cm. To solve the problem more formally with a conversion factor, we first write the quantity we are given, 3.55 m. Then we multiply this quantity by a conversion factor, which is the same as multiplying it by 1. We can write 1 as $\frac{100 \text{ cm}}{1 \text{ m}}$ and multiply:

$$3.55 \text{ m} \times \frac{100 \text{ cm}}{1 \text{ m}} \tag{3.1.5}$$

The 3.55 m can be thought of as a fraction with a 1 in the denominator. Because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out:

$$\frac{3.55 \ \cancel{\text{m}}}{1} \times \frac{100 \text{ cm}}{1 \ \cancel{\text{m}}} \tag{3.1.6}$$

The final step is to perform the calculation that remains once the units have been canceled:

$$\frac{3.55}{1} \times \frac{100 \text{ cm}}{1} = 355 \text{ cm} \tag{3.1.7}$$

In the final answer, we omit the 1 in the denominator. Thus, by a more formal procedure, we find that 3.55 m equals 355 cm. A generalized description of this process is as follows:

quantity (in old units) × conversion factor = quantity (in new units)

You may be wondering why we use a seemingly complicated procedure for a straightforward conversion. The conversion problems you encounter *will not always be so simple*. If you master the technique of applying conversion factors, you will be able to solve a large variety of problems.

In the previous example, we used the fraction $\frac{100 \text{ cm}}{1 \text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1 \text{ m}}{100 \text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

$$3.55 \text{ m} \times \frac{1 \text{ m}}{100 \text{ cm}} = 0.0355 \frac{\text{m}^2}{\text{cm}} \tag{3.1.8}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 3.1.1 shows a concept map for constructing a proper conversion.
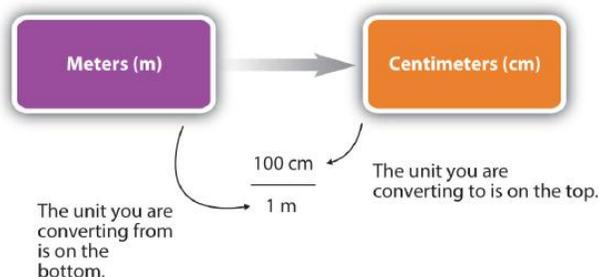

Figure 3.1.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another.

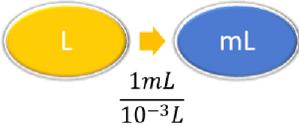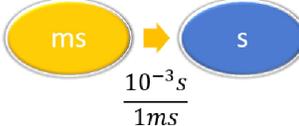📌 **General Steps in Performing Dimensional Analysis**

1. Identify the "**given**" information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to "**find**"? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

- Numbers in conversion factors based on SI definitions, such as kilograms to grams, are *not* considered in the determination of significant figures in a calculation because the numbers in such conversion factors are exact.
- Conversions within the English system are also exact. The numbers in conversion factors from one English unit to another English unit do not effect the significant figures in the answer.
- Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact.
- In contrast, conversion factors that come from measurements (such as density, as we will see shortly) have a limited number of significant figures and should be considered in determining the significant figures of the final answer.
- Conversions between the SI and English systems are not exact. The significant figures in the conversion must be considered when determining the significant figures in the answer. The one exception is the conversion between inches and centimeters (1 in = 2.54 cm) which is exact.

✔ **Example 3.1.1**

Steps for Problem Solving	Example 3.1.1 The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	Example 3.1.2 A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\,mL = 10^{-3}L$	$1\,ms = 10^{-3}s$
Prepare a concept map and use the proper conversion factor.	L ➡ mL $\dfrac{1mL}{10^{-3}L}$	ms ➡ s $\dfrac{10^{-3}s}{1ms}$
Cancel units and calculate.	$4.7\,\cancel{L} \times \dfrac{1\,\text{mL}}{10^{-3}\,\cancel{L}} = 4,700\,\text{mL}$ or $4.7\,\cancel{L} \times \dfrac{1,000\,\text{mL}}{1\,\cancel{L}} = 4,700\,\text{mL}$ or 4.7 x 10^3 mL 2SF, not ambiguous	$18\,\cancel{ms} \times \dfrac{10^{-3}\,\text{s}}{1\,\cancel{ms}} = 0.018\,\text{s}$ or $18\,\cancel{ms} \times \dfrac{1\,\text{s}}{1,000\,\cancel{ms}} = 0.018\,\text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

> **?** **Exercise 3.1.1**
>
> Perform each conversion.
>
> a. 101,000 ns to seconds
> b. 32.08 kg to grams
> c. 1.53 grams to cg
>
> **Answer a:**
> $1.01 x 10^{-4} s$
> **Answer b:**
> $3.208 x 10^{4} g$
> **Answer c:**
> $1.53 x 10^{2} cg$

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3.2: Multi-Step Conversion Problems

> **◐ Learning Objectives**
>
> - Given a quantity, convert from one set of units to another using dimensional analysis showing canceling on units. This includes multiple factor conversions.

Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence. We first convert the given amount in km to the base unit, which is meters. We know that 1,000 m =1 km.

Then we convert meters to mm, remembering that $1 \text{ mm} = 10^{-3}$ m.

Concept Map
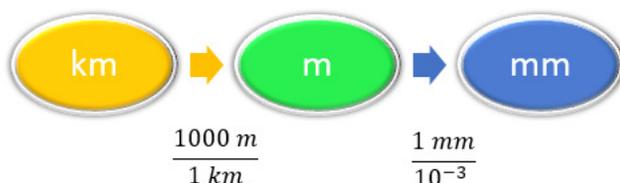
$$\frac{1000 \, m}{1 \, km} \qquad \frac{1 \, mm}{10^{-3}}$$

Calculation

$$54.7 \ \cancel{km} \times \frac{1,000 \ \cancel{m}}{1 \ \cancel{km}} \times \frac{1 \ mm}{10^{-3} \ \cancel{m}} = 54,700,000 \ mm$$

$$= 5.47 \times 10^{7} \ mm$$

> *In each step, the previous unit is canceled and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.*

> **✔ Example 3.2.1: Unit Conversion**
>
> Convert 58.2 ms to megaseconds in one multi-step calculation.
>
> **Solution**
>
> | Steps for Problem Solving | Unit Conversion |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given: 58.2 ms
 Find: Ms |
> | List other known quantities | $1ms = 10^{-3}s$
 $1Ms = 10^{6}s$ |
> | Prepare a concept map. |
 $$\frac{10^{-3}s}{1 \, ms} \qquad \frac{1 \, Ms}{10^{6} \, s}$$ |
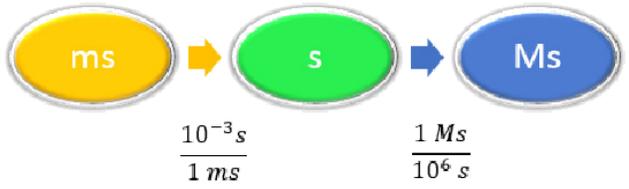
Steps for Problem Solving	Unit Conversion
Calculate.	$58.2 \ \cancel{ms} \times \dfrac{10^{-3} \ \cancel{s}}{1 \ \cancel{ms}} \times \dfrac{1 \ Ms}{1,000,000 \ \cancel{s}} = 0.0000000582 \ Ms$ $= 5.82 \times 10^{-8} \ Ms$ Neither conversion factor affects the number of significant figures in the final answer.

✔ Example 3.2.2: Unit Conversion

How many seconds are in 2.50 days?

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 2.50 days Find: s
List other known quantities.	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds
Prepare a concept map.	 day → hr → min → s $\dfrac{24hr}{1 \ day}$ $\dfrac{60min}{1 \ hr}$ $\dfrac{60 \ s}{1 \ min}$
Calculate.	$2.50 \ d \times \dfrac{24 \ hr}{1 \ d} \times \dfrac{60 \ min}{1 \ hr} \times \dfrac{60 \ s}{1 \ min} = 216,000 \ s$ 2.16×10^5 s 3SF, not ambiguous
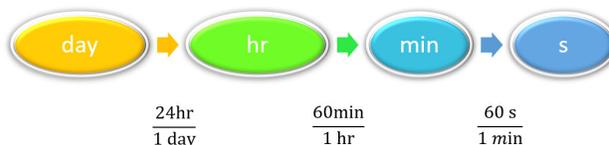
? Exercise 3.2.1

Perform each conversion in one multi-step calculation.

 a. 43.007 ng to kg
 b. 1005 in to yd

Answer a
  $4.3007 \times 10^{-11} \, kg$

Answer b
  $27.92 \, yd$

Contributions

- Marisa Alviar-Agnew (Sacramento City College)

- Henry Agnew (UC Davis)

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3.3: Units Raised to a Power

> **Learning Objectives**
>
> - Given a quantity, convert from one set of units to another using dimensional analysis showing canceling on units. This includes numbers and units raised to a power.

Conversion factors for area and volume can also be produced by the dimensional analysis method. Just remember that if a quantity is raised to a power of 10, both the number and the unit must be raised to the same power of 10. For example, to convert 1500 cm^2 to m^2, we need to start with the relationship between centimeter and meter. We know that 1 cm $= 10^{-2}$ m or 100 cm $=1$ m, but since we are given the quantity in 1500 cm^2, then we have to use the relationship:

$$1\ cm^2 = (10^{-2}\ m)^2 = 10^{-4}\ m^2 \tag{3.3.1}$$

CONCEPT MAP
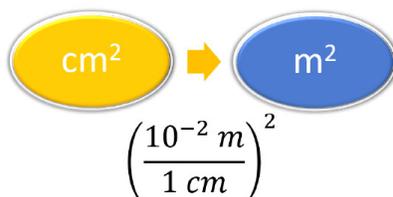
$$\left(\frac{10^{-2}\ m}{1\ cm}\right)^2$$

CALCULATION

$$1500\ \cancel{cm}^2 \times \left(\frac{10^{-2}\ m}{1\ \cancel{cm}}\right)^2 = 0.15\ m^2 \tag{3.3.2}$$

or

$$1500\ \cancel{cm}^2 \times \left(\frac{1\ m}{100\ \cancel{cm}}\right)^2 = 0.15\ m^2 \tag{3.3.3}$$

or

$$1500\ \cancel{cm}^2 \times \frac{1\ m^2}{10,000\ \cancel{cm}^2} = 0.15\ m^2 \tag{3.3.4}$$

> **✔ Example 3.3.1: Unit Conversion**
>
> The volume of a sphere is 333 in^3. What is the volume in cubic cm (cm^3)?
>
> **Solution**
>
> | Steps for Problem Solving | What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm^3)? |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given: 333 in^3
 Find: cm^3 |
> | Determine other known quantities. | 1 in^3 = (2.54 cm)^3
 1 in^3 = 16.4 cm^3 |
> | Prepare a concept map. | $\left(\frac{2.54\ cm}{1\ in}\right)^3$ |
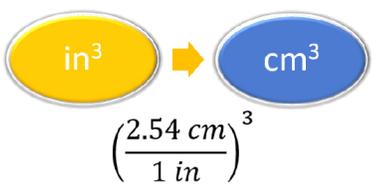
Steps for Problem Solving	What is the volume of a sphere (radius 4.30 inches) in cubic cm (cm³)?
Calculate.	$333\ in^3\left(\dfrac{2.54cm}{1\ in}\right)^3 = 5.46\times 10^3\,cm^3$
Think about your result.	A centimeter is a smaller unit than an inch, so the answer in cubic centimeters is larger than the given value in cubic inches.

? Exercise 3.3.1

Lake Tahoe has a surface area of 191 square miles. What is the area in square km (km²)?

Answer

495 km²

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3.4: Units in the Numerator and the Denominator

> **◗ Learning Objectives**
>
> - Given a quantity, convert from one set of units to another using dimensional analysis showing canceling on units. This includes compound units (eg: mph to kps).

Compound Units

There are many units that combine a unit in the numerator with a unit in the denominator. Common examples are miles per hour, meters per second, and miles per gallon. The "per" that appears when we talk about the compound unit or when we write it in text form translates to the divider between numerator and denominator when we write it symbolically. Therefore miles per hour would be written mi/h, meters per second would be written m/s, and miles per gallon would be written mi/gal.

For example, suppose we want to convert 0.0128 kilometers per minute to meters per hour.

Concept Map
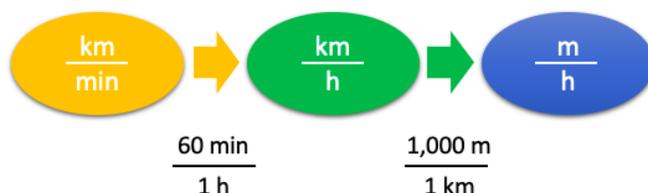
Calculation

$$\frac{0.0128 \ \cancel{km}}{1 \ \cancel{min}} \times \frac{60 \ \cancel{min}}{1 \ h} \times \frac{1,000 \ m}{1 \ \cancel{km}} = \frac{768 \ m}{h}$$
$$= 768 \ m/h$$

Which we would write or say as 768 meters per hour.

> **✔ Example 3.4.1: Unit Conversion**
>
> Convert 45 miles per gallon to kilometers per liter in one multi-step calculation.
>
> **Solution**
>
> | Steps for Problem Solving | Unit Conversion |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given: 45 mi/gal
 Find: km/L |
> | List other known quantities | $1 km = 0.6214 mi$
 $1 gal = 3.785 L$ |
> | Prepare a concept map. | |

Steps for Problem Solving	Unit Conversion
Calculate.	$$\frac{45 \ \cancel{mi}}{1 \ \cancel{gal}} \times \frac{1 \ \cancel{gal}}{3.785 \ L} \times \frac{1 \ km}{0.6214 \ \cancel{mi}} = \frac{19 \ km}{L}$$ $$= 19 \ km/L$$

Convert 1.25 cm/s to in/min.

Answer

$29.5 \ in/min$

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3.5: Density

> **◐ Learning Objectives**
>
> - Given a mass and volume calculate the density of a substance.

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 is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \tag{3.5.1}$$

or just

$$\rho = \frac{m}{V} \tag{3.5.2}$$

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 3.5.1. 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 3.5.1: 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 or g/mL. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

> **✔ Example 3.5.1: 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 3.5.2:
>
> $$\rho = \frac{23.71002\,g}{30.2\,mL} = 0.785\,g/mL$$

> **? Exercise 3.5.1**
>
> 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:

$$2.0 \, \cancel{mL} \times \frac{13.6 \text{ g}}{1 \, \cancel{mL}} = 27.2 \text{ g} = 27 \text{ g}$$
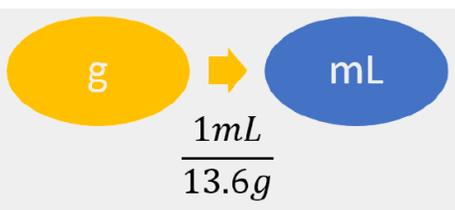
In the last step, we limit our final answer to two significant figures because the volume quantity has only two significant figures; the 1 in the volume unit is considered an exact number, so it does not affect the number of significant figures. The other conversion factor would be useful if we were 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 3.5.2: 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

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.	$\dfrac{1mL}{13.6g}$
Calculate.	$0.750 \, \cancel{g} \times \dfrac{1 \text{ mL}}{13.6 \, \cancel{g}} = 0.0551 \text{ mL} \qquad (3.5.3)$

Density is a measurement, and like any measurement it contains some error. The significant figures in the density must therefore be considered when determining the significant figures in any answer calculated using it. Consider, for example, if we had tried

calculating the volume of 0.7500 g of mercury in the previous example. Although the mass we are starting with now has four significant figures, the density we are using as a conversion factor still only has three significant figures. We would therefore still need to report the answer to three significant figures, 0.0551 mL.

> **? Exercise 3.5.2**
>
> What is the volume of 100.0 g of air if its density is 1.3 g/L?
>
> **Answer**
> $77 \, L$

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3.6: Temperature

> **⬤ Learning Objectives**
>
> - Convert temperatures between Celsius, Kelvin, and Fahrenheit scales.

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter. As temperature increases, the kinetic energy of the particles also increases. **Heat**, on the other hand, is the flow of thermal energy from an object with a higher temperature to an object with a lower temperature. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is almost always because there was an 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 3.6.1: 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.

The Fahrenheit Scale

The main problem with the Fahrenheit scale is the arbitrary definitions of temperature. The freezing point of water was defined as 32^oF and the boiling point as 212^oF. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

The Celsius scale sets the freezing point and boiling point of water at 0^oC and 100^oC 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 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.
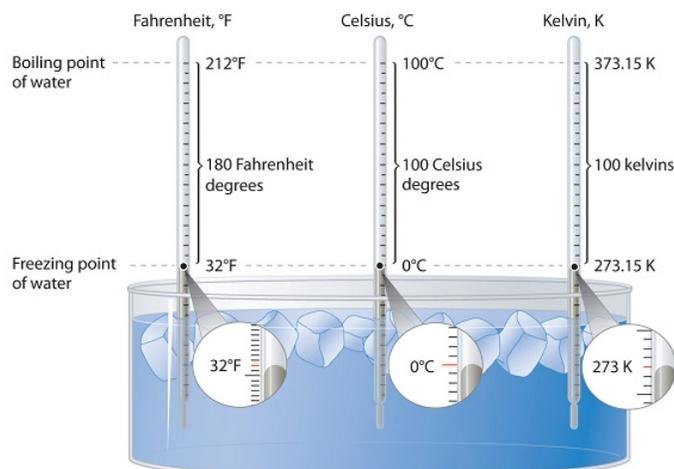
Figure 3.6.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In 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:

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

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

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:

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

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

✔ **Example 3.6.1: 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:

$$^\circ C = \frac{(103.5\,^\circ F - 32)}{1.8} \tag{3.6.1}$$

$$= 39.7\,^\circ C \tag{3.6.2}$$

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

$$K = 39.7\,^\circ C + 273.15 \tag{3.6.3}$$

$$= 312.9\,K \tag{3.6.4}$$

? **Exercise 3.6.1**

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 °C, 9981 °F

Answer (b)

    2807 °C, 5085 °F

Answer (c)

    -195.79 °C, -320.42 °F

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CHAPTER OVERVIEW

4: Classifiation of Matter- Properties and Changes

4.1: Introduction to the Nanoscopic Level- Solids, Liquids, and Gases

Learning Objectives

- To describe the solid, liquid and gas phases.

Water can take many forms. At low temperatures (below 0°C), it is a solid. When at "normal" temperatures (between 0°C and 100°C), it is a liquid. While at temperatures above 100°C, water is a gas (steam). The state that water is in depends upon the temperature. Each state has its own unique set of physical properties. Matter typically exists in one of three states: **solid**, **liquid**, or **gas**.

Figure 4.1.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. Most substances can exist in any of these three states. Figure 4.1.2 shows the differences among solids, liquids, and gases. On the macroscopic level, a solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.
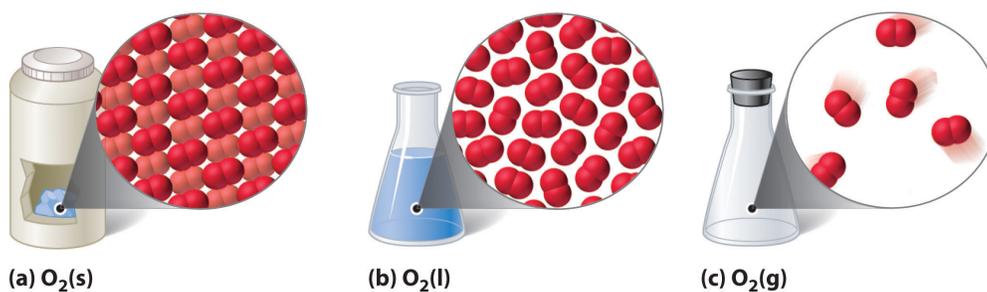

(a) O$_2$(s)   **(b) O$_2$(l)**   **(c) O$_2$(g)**

Figure 4.1.2: A Representation of the Solid, Liquid, and Gas States. (a) Solid O$_2$ has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O$_2$ conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O$_2$ fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

📌 Plasma: A Fourth State of Matter

Technically speaking, a fourth state of matter called plasma exists, but it does not naturally occur on earth, so we will omit it from our study here.
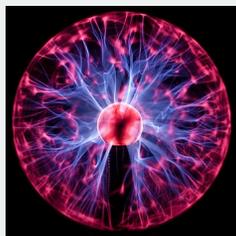
A plasma globe operating in a darkened room. (CC BY-SA 3.0; Chocolateoak).

Solids

In the solid state, the individual particles of a substance are in fixed positions with respect to each other because there is not enough thermal energy to overcome the attractive forces between the particles (known as intermolecular forces). As a result, solids have a definite shape and volume. Most solids are hard, but some (like waxes) are relatively soft. Many solids composed of ions can also be quite brittle.

Solids are defined by the following characteristics:

- Definite shape (rigid)
- Definite volume
- Particles vibrate around fixed axes

If we were to cool liquid mercury to its freezing point of -39°C, and under the right pressure conditions, we would notice all of the liquid particles would go into the solid state. Mercury can be solidified when its temperature is brought to its freezing point. However, when returned to room temperature conditions, mercury does not exist in solid state for long, and returns back to its more common liquid form.

Solids often have their constituent particles arranged in a regular, three-dimensional array called a **crystal**, as shown in Figure 4.1.3. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, "without form") solids. Glass is one example of an amorphous solid.
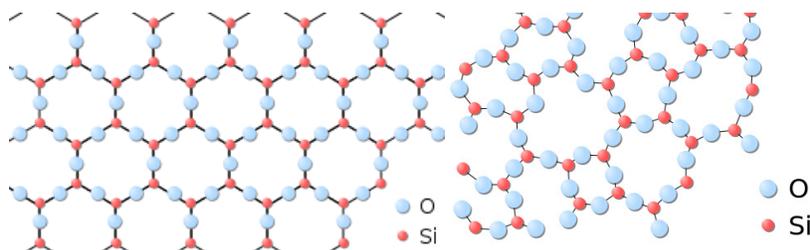

Figure $4.1.3$: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

If the particles of a substance have enough energy to partially overcome intermolecular interactions, then the particles can move about each other while remaining in contact. This describes the liquid state. In a liquid, the particles are still in close contact, so liquids have a definite volume. However, because the particles can move about each other rather freely, a liquid has no definite shape and takes a shape dictated by its container.

Liquids have the following characteristics:

- No definite shape (takes the shape of its container).
- Has definite volume.
- Particles are free to move over each other, but are still attracted to each other.

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension)—a property that all liquids exhibit. Mercury has a relatively high surface tension, which makes it very unique. Here you see mercury in its common liquid form.

Video **4.1.1**: *Mercury boiling to become a gas.*

If we heat liquid mercury to its boiling point of 357^oC under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. This describes the gas state. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either. The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Gases have the following characteristics:

- No definite shape (takes the shape of its container)
- No definite volume (another way of saying this is that a gas is highly compressible)
- Particles move in random motion with little or no attraction to each other

Table 4.1.1: Characteristics of the Three States of Matter

Characteristics	Solids	Liquids	Gases
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
strength of intermolecular forces relative to kinetic energy	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

> ✔ **Example 4.1.1**
>
> What state or states of matter does each statement, describe?
>
> a. This state has a definite volume, but no definite shape.
> b. This state has no definite volume.
> c. This state allows the individual particles to move about while remaining in contact.
>
> **Solution**
>
> a. This statement describes the liquid state.
> b. This statement describes the gas state.
> c. This statement describes the liquid state.

? **Exercise 4.1.1**

What state or states of matter does each statement describe?

  a. This state has individual particles in a fixed position with regard to each other.
  b. This state has individual particles far apart from each other in space.
  c. This state has a definite shape.

Answer a:
    solid

Answer b:
    gas

Answer c:
    solid

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4.2: Classifying Matter- the Macroscopic Level

Learning Objectives

- Given a macroscopic definition classify the type of matter.

Matter can be classified using a hierarchy (Figure 4.2.1). In this section, we will think about this hierarchy in terms of macroscopic properties. In the next section, we will revisit this hierarchy based on nanoscopic structure.
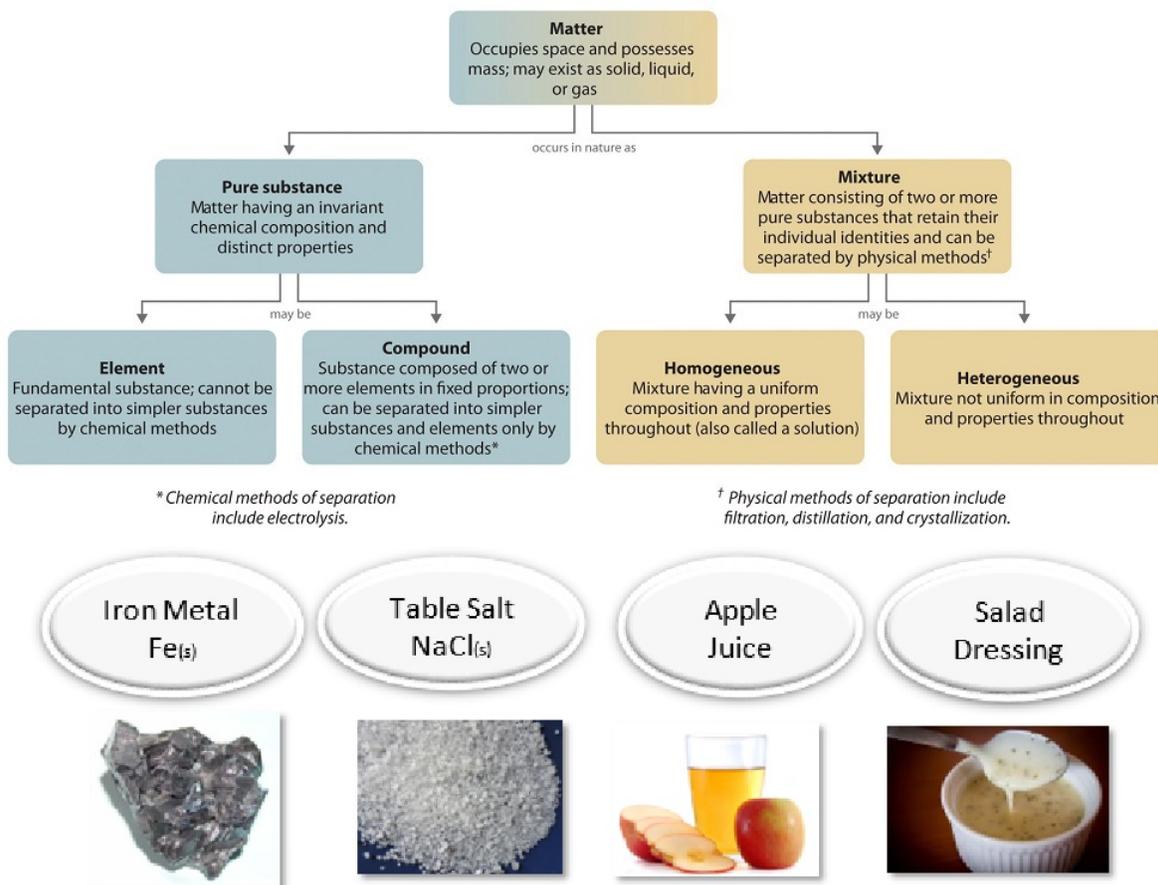
Figure 4.2.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

A **pure substance** always has the same composition and same physical properties regardless of the source of the sample. For example, pure water is always composed of two atoms of hydrogen for every one atom of oxygen. This means its composition is always the same. Also, the physical properties of pure water will always be the same: at "normal" pressures it will boil at $100^{o}C$, freeze at $0^{o}C$, and have a density of 1.00 g/ml.

Contrast this with salt water. The amount of salt in the water can vary, meaning the composition of salt water is not always the same. Additionally, the density, freezing print, and boiling point will change depending how much salt is in the water. Variable composition and variable properties means that salt water is not a pure substance. In fact, it is composed of two pure substances: water and sodium chloride. Any matter composed of multiple pure substances is a **mixture**. Mixtures can be separated into pure substances by means of physical separations.

There are two types of pure substances: compounds and elements. **Compounds** can be broken into simpler pure substances by chemical means. When a current with a sufficient voltage is passed through water, for example, it can be broken into oxygen and hydrogen. This process is called hydrolysis and is one example of a chemical separation. Since water is a pure substance that can be broken into simpler pure substances by chemical means, water is a compound.

A pure substance that can not be broken down into simpler substances by chemical means is an element. You might already know from a previous science course that the periodic table contains all the known elements. The oxygen and hydrogen that form when

water undergoes hydrolysis are examples of elements. They cannot be broken down into simpler substances by any means. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

Mixtures can also be further divided: some are homogeneous and some are heterogeneous.

A **homogeneous mixture** is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. The amount of salt in the salt water can vary from one sample to another. All solutions are considered homogeneous because the dissolved material is present in the same amount throughout the solution.

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

📌 **Phase**

A phase is any part of a sample that has a uniform composition and properties. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly, but instead form two separate layers. Each of the layers is called a phase.

✔ **Example 4.2.1**

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

a. filtered tea
b. freshly squeezed orange juice
c. a compact disc
d. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
e. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
   B) Because the composition of the solution is uniform throughout, it is *a homogeneous mixture*.
b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
   B) Because its composition is not uniform throughout, orange juice is *a heterogeneous mixture*.
c. **A)** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence, a compact disc is not chemically pure.
   B) The regions of different composition indicate that a compact disc is *a heterogeneous mixture.*
d. **A)** Aluminum oxide is a single, chemically *pure compound*.
e. **A)** Selenium is one of the known *elements*.

? Exercise 4.2.1

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

a. white wine
b. mercury
c. ranch-style salad dressing
d. table sugar (sucrose)

Answer a:
　homogeneous mixture (solution)
Answer b:
　element
Answer c:
　heterogeneous mixture
Answer d:
　compound

✔ Example 4.2.2

How would a chemist categorize each example of matter?

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

Solution

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.
b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
c. Water is a substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a substance, is an element.

? Exercise 4.2.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a:
　a homogeneous mixture (solution), assuming it is filtered coffee
Answer b:
　element
Answer c:
　heterogeneous mixture

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4.3: Classifying Matter- the Nanoscopic Level

◐ Learning Objectives

- Given a particle level picture or chemical formula classify the type of matter.

Particles

In the previous section, matter was classified based on macroscopic observations such as the ability of a sample to be separated into simpler substances or whether the appearance of the sample is consistent throughout all its parts. These macroscopic observations are manifestations of the nanoscopic structure. In this section, different shapes will be used to represent atoms of different elements. We will use the term **particle** to refer to an atom or group of atoms that are chemically bonded together.

Pure Substances

The drawings below all represent pure substances on the nanoscopic level.
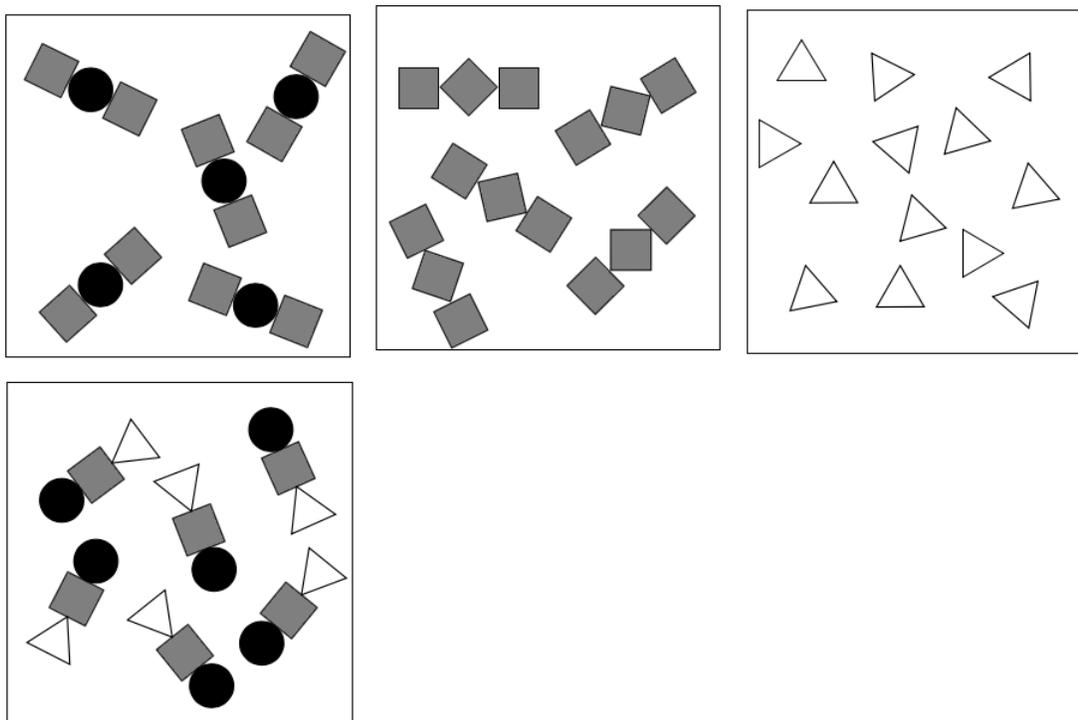
Figure 4.3.1: Nanoscopic Level Representations of Pure Substances

In a pure substance, all the particles are the same. Notice that the particle may be composed of a single atom or multiple atoms. If there are multiple atoms, the atoms may be all the same element or they may be different elements. The important feature is that the same particle is repeated over and over again throughout the sample.

Mixtures

Notice how the drawings below, which all represent mixtures, differ from the drawings of pure substances above.
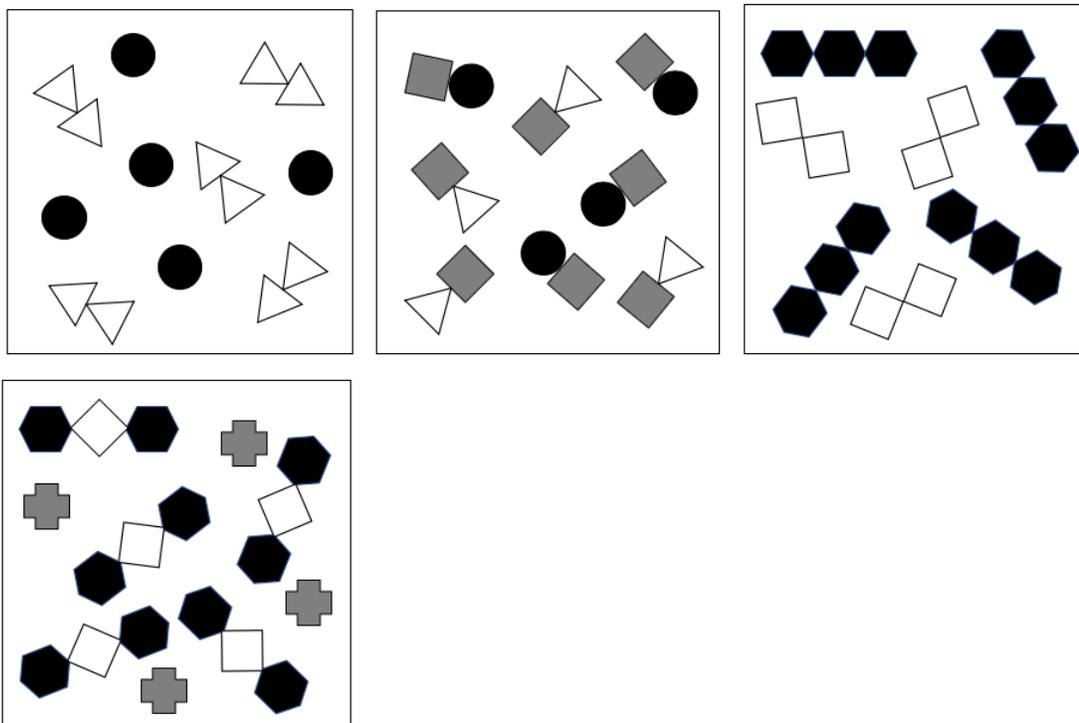
Figure 4.3.2: Nanoscopic Level Representations of Mixtures

In a mixture, all the particles are not the same. That is to say, a mixture is composed of two or more pure substances. In homogeneous mixtures, the composition of the particles is the same throughout. In heterogeneous mixtures, particular regions are richer in one type of particle relative to other regions in the sample.

Elements

Recall that pure substances can be further subdivided into elements and compounds. The following drawings all represent sample of elements.
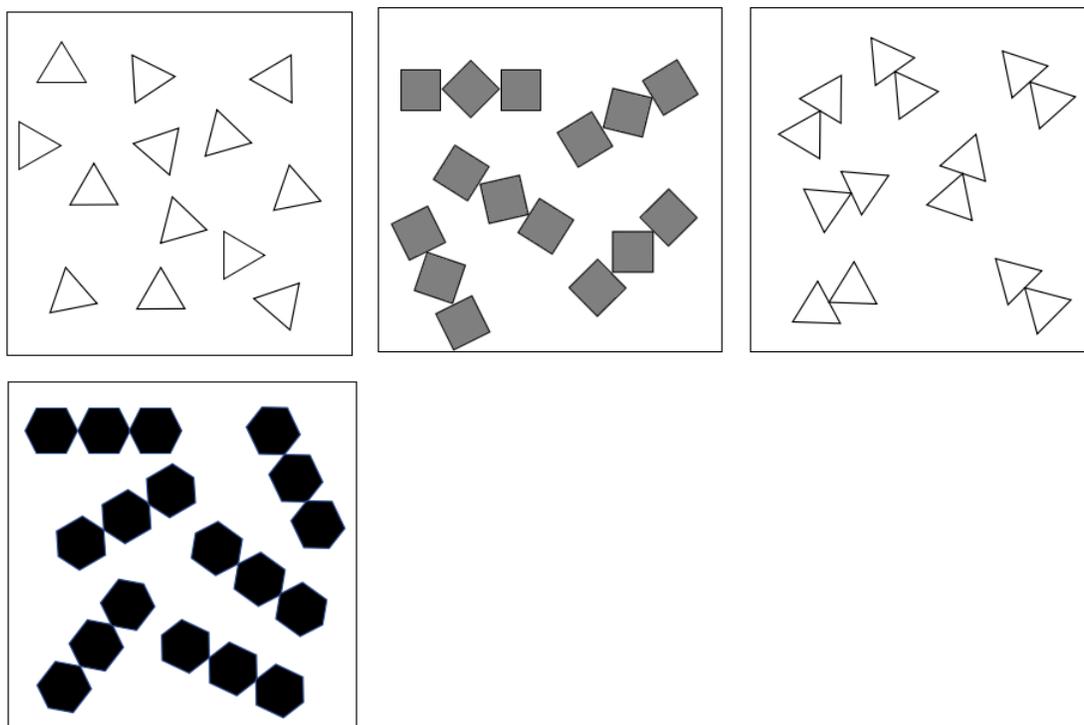
Figure $4.3.3$: Nanoscopic Level Representations of Elements

In a sample of an element, not only are all the particles the same, but all of the atoms that compose the particle are the same. Notice that the particle may contain a single atom or multiple atoms.

Compounds

Compound might be the most important chemistry term that is routinely misunderstood by beginning chemistry students. Mistakenly, mixtures are often misidentified as compounds. Play close attention to the drawings of compounds below.
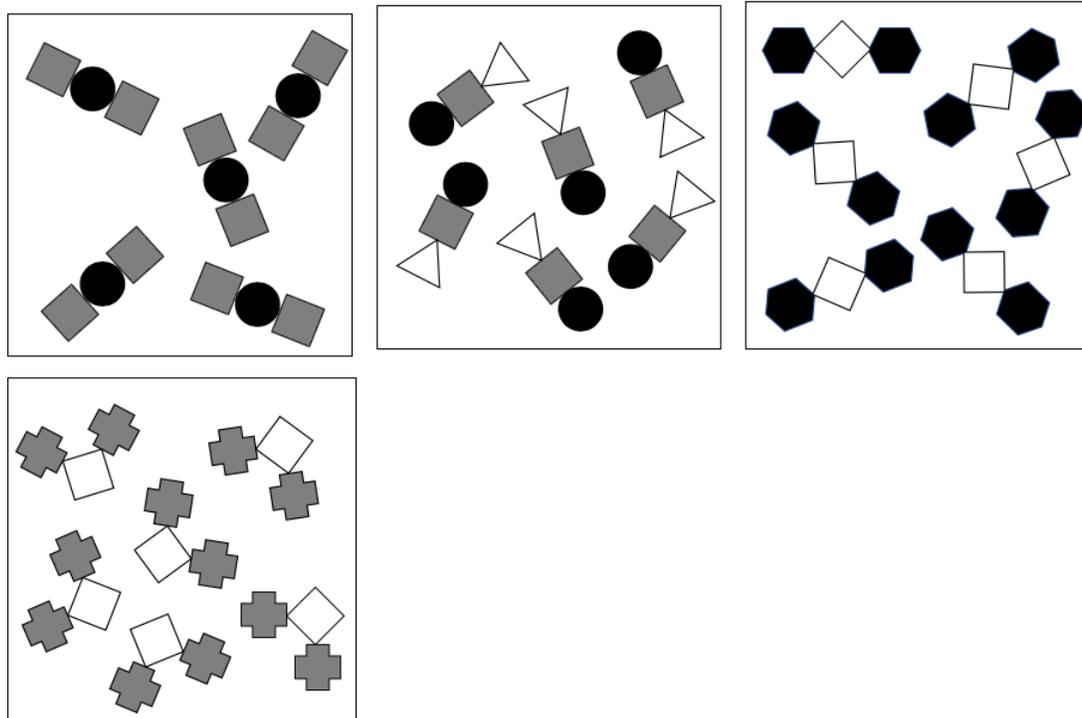
Figure 4.3.4: Nanoscopic Level Representations of Compounds

Two requirements must be met for a sample of matter to be called a compound. First, all of the particles must be the same. This is another way of saying that a compound is a pure substance. Second, the particle in a compound must include at least two different types of atoms.

Chemical Formulas

A chemical formula is a symbolic way of representing the atoms that compose a particle. A formula indicates both the elements present in a particle and the number of atoms of each element. We cannot write a chemical formula for a mixture since there are a multiple particles with different compositions present.

Chemical formulas consist of the symbols for the elements present, and each symbol is followed by a subscript. Symbols for elements are either a single upper case letter or an upper case letter followed by a lower case letter. Subscripts (small numbers written to the right of the symbol and sitting slightly lower than the symbol) indicate the number of that type of atom in the particle. Subscript of one are not written, so that a missing subscript is understood to mean there is only one of a particular type of atom in the particle.

Consider the following three fictitious elements and their symbols:

element	●	■	△
symbol	Z	T	X

These symbols can be used to write the chemical formulas for some of the pure substances represented by particle-level drawings in this section:
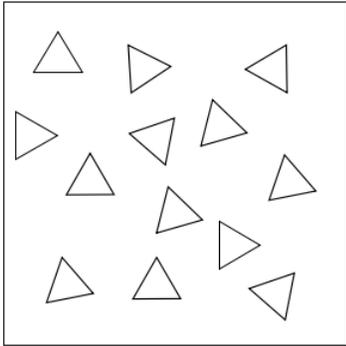
particle-level drawing	chemical formula	type of pure substance
	X	element
	X_2	element
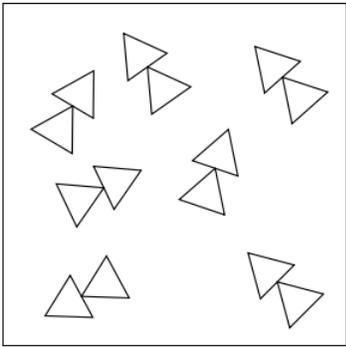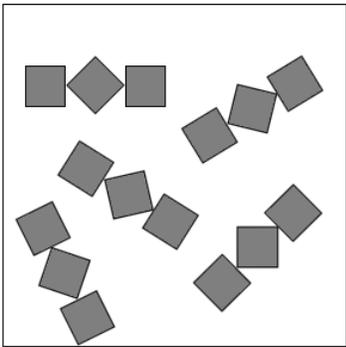
particle-level drawing	chemical formula	type of pure substance
	T$_3$	element
	ZTX	compound
	ZT$_2$	compound
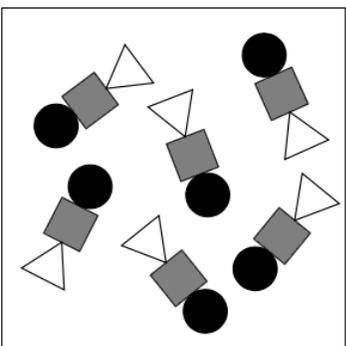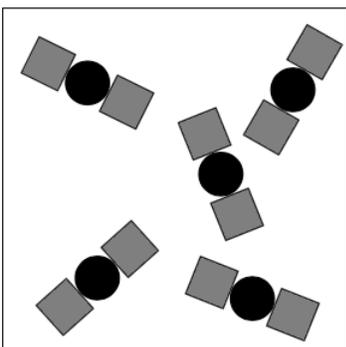
4.4: The Elements

> **◑ Learning Objectives**
>
> - Define a chemical element.
> - Represent a chemical element with a chemical symbol.

An element is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. Figure 4.4.1 shows some of the chemical elements.

Figure 4.4.1: Samples of Elements. Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature. © Thinkstock

Abundance

The elements vary widely in abundance. In the universe as a whole, the most common element is hydrogen (about 90% of atoms), followed by helium (most of the remaining 10%). All other elements are present in relatively minuscule amounts, as far as we can detect.

Table 4.4.1: Elemental Composition of Earth

Earth's Crust		Earth (overall)	
Element	**Percentage**	**Element**	**Percentage**
oxygen	46.1	iron	34.6
silicon	28.2	oxygen	29.5
aluminum	8.23	silicon	15.2
iron	5.53	magnesium	12.7
calcium	4.15	nickel	2.4
sodium	2.36	sulfur	1.9
magnesium	2.33	all others	3.7
potassium	2.09		
titanium	0.565		
hydrogen	0.14		
phosphorus	0.105		
all others	0.174		
Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 14–17.			

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.2%. Hydrogen,

the most abundant element in the universe, makes up only 0.14% of Earth's crust. Table 4.4.1 lists the relative abundances of elements on Earth as a whole and in Earth's crust. Table 4.4.2 lists the relative abundances of elements in the human body. If you compare Table 4.4.1 and Table 4.4.2, you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in Table 4.4.1; carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table 4.4.2: Elemental Composition of a Human Body

Element	Percentage by Mass
oxygen	61
carbon	23
hydrogen	10
nitrogen	2.6
calcium	1.4
phosphorus	1.1
sulfur	0.20
potassium	0.20
sodium	0.14
chlorine	0.12
magnesium	0.027
silicon	0.026
iron	0.006
fluorine	0.0037
zinc	0.0033
all others	0.174

Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 7–24.

The relative amounts of elements in the body have less to do with their abundances on Earth than with their availability in a form we can assimilate. We obtain oxygen from the air we breathe and the water we drink. We also obtain hydrogen from water. On the other hand, although carbon is present in the atmosphere as carbon dioxide, and about 80% of the atmosphere is nitrogen, we obtain those two elements from the food we eat, not the air we breathe.

Names and Symbols

Each element has a name. Some of these names date from antiquity, while others are quite new. Today, the names for new elements are proposed by their discoverers but must be approved by the International Union of Pure and Applied Chemistry, an international organization that makes recommendations concerning all kinds of chemical terminology.

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter chemical symbol. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. Table 4.4.3 lists the names and symbols of some of the most familiar elements.

Table 4.4.3: Element Names and Symbols

aluminum	Al	magnesium	Mg
argon	Ar	manganese	Mn
arsenic	As	mercury	Hg*
barium	Ba	neon	Ne
bismuth	Bi	nickel	Ni
boron	B	nitrogen	N
bromine	Br	oxygen	O
calcium	Ca	phosphorus	P
carbon	C	platinum	Pt
chlorine	Cl	potassium	K*
chromium	Cr	silicon	Si
copper	Cu*	silver	Ag*
fluorine	F	sodium	Na*
gold	Au*	strontium	Sr
helium	He	sulfur	S
hydrogen	H	tin	Sn*
iron	Fe	tungsten	W†
iodine	I	uranium	U
lead	Pb*	zinc	Zn
lithium	Li	zirconium	Zr

*The symbol comes from the Latin name of element.

†The symbol for tungsten comes from its German name—*wolfram*.

Symbols for elements that seem odd for English speakers in many cases make sense to speakers of languages such as Spanish and French that are descended from Latin. For example, gold is *oro* in Spanish and *or* in French (close to the Latin *aurum*), tin is *estaño* in Spanish (compare to *stannum*), lead is *plomo* in Spanish and *plomb* in French (compare to *plumbum*), silver is *argent* in French (compare to *argentum*), and iron is *fer* in French and *hierro* in Spanish (compare to *ferrum*). The closeness is even more apparent in pronunciation than in spelling.

✔ Example 4.4.1

Write the chemical symbol for each element without consulting Table 4.4.3 "Element Names and Symbols".

a. bromine
b. boron
c. carbon
d. calcium
e. gold

Answer a

Br

Answer b

B

Answer c

C

Answer d

Ca

Answer e

Au

? Exercise 4.4.1

Write the chemical symbol for each element without consulting Table 4.4.3.

a. manganese
b. magnesium
c. neon
d. nitrogen
e. silver

Answer a

Mn

Answer b

Mg

Answer c

Ne

Answer d

N

Answer e

Ag

✔ Example 4.4.2

What element is represented by each chemical symbol?

a. Na
b. Hg
c. P
d. K
e. I

Answer a

sodium

Answer b

mercury

Answer c

phosphorus

Answer d

potassium

Answer e

    iodine

? Exercise 4.4.2

What element is represented by each chemical symbol?

  a. Pb
  b. Sn
  c. U
  d. O
  e. F

Answer a

    lead

Answer b

    tin

Answer c

    uranium

Answer d

    oxygen

Answer e

    fluorine

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4.5: Physical and Chemical Changes- the Macroscopic Level

> **◑ Learning Objectives**
>
> - Recognize physical evidence that a chemical reaction has occurred.

Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are classified as either physical or chemical changes. Chemists learn a lot about the nature of matter by studying the changes that matter can undergo. Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.

Physical Change

Physical changes are changes in which no chemical bonds are broken or formed. This means that the same types of compounds or elements that were there at the beginning of the change are there at the end of the change. Physical changes involve moving molecules around, but not changing them. Some types of physical changes include:

- Changes of state (changes from a solid to a liquid or a gas and vice versa).
- Separation of a mixture.
- Physical deformation (cutting, denting, stretching).
- Making solutions (dissolving one substance in another).

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. When liquid water is heated, it changes to water vapor. However, even though the physical properties have changed, the molecules are exactly the same as before. We still have each water molecule containing two hydrogen atoms and one oxygen atom covalently bonded. When you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of the pennies or the nickels—you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you started tearing is still paper when you are done. Again, this is an example of a physical change.

Figure 4.5.1: 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 composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

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.

Chemical Change

Chemical changes occur when bonds are broken and/or formed between molecules or atoms. This means that one substance with a certain set of properties (such as melting point, color, taste, etc) is turned into a different substance with different properties. Chemical changes are frequently harder to reverse than physical changes.

One good example of a chemical change is burning a candle. The act of burning a candle results in the formation of new compounds (carbon dioxide and water) from the burning of the wax and oxygen in the air. Not only has the appearance changes, but the structure of the molecules have also changed. The new substances do not have the same chemical properties as the original ones.

Figure 4.5.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

We can't actually see molecules breaking and forming bonds, although that's what defines chemical changes. We have to make other observations to indicate that a chemical change has happened. Some of the evidence for chemical change will involve the energy changes that occur in chemical changes, but some evidence involves the fact that new substances with different properties are formed in a chemical change.

Observations that help to indicate chemical change include:

- Temperature changes (either the temperature increases or decreases).
- Light given off.
- Unexpected color changes (a substance with a different color is made, rather than just mixing the original colors together).
- Bubbles are formed (but the substance is not boiling—you made a substance that is a gas at the temperature of the beginning materials, instead of a liquid).
- Different smell or taste (do not taste your chemistry experiments, though!).
- A solid forms if two clear liquids are mixed (look for *floaties*—technically called a precipitate).

> ✔ **Example 4.5.1**
>
> Label each of the following changes as a physical or chemical change. Give evidence to support your answer.
>
> a. Boiling water.
> b. A nail rusting.
> c. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
> d. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.
>
> **Solution**
>
> a. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $H_2O\,(l) \rightarrow H_2O\,(g)$
> b. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $Fe + O_2 \rightarrow Fe_2O_3$
> c. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
> d. Chemical: the formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

> **?** Exercise 4.5.1
>
> Label each of the following changes as a physical or chemical change.
>
> a. A mirror is broken.
> b. An iron nail corroded in moist air
> c. Copper metal is melted.
> d. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.
>
> **Answer a:**
>
> physical change
>
> **Answer b:**
>
> chemical change
>
> **Answer c:**
>
> physical change
>
> **Answer d:**
>
> chemical change

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4.6: Physcial and Chemical Changes- the Nanoscopic Level

> **◐ Learning Objectives**
>
> - Classify changes in matter as physical or chemical on the particle level.

Physical Changes

Chemical and physical changes can also be understood using nanoscopic level illustrations. Representations of three physical changes are shown below. In all of these illustrations, the arrow represents the process of transformation. What is to the left of the arrow can be thought of as a "before" snapshot and what is to the right of the arrow can be thought of as an "after" snapshot.
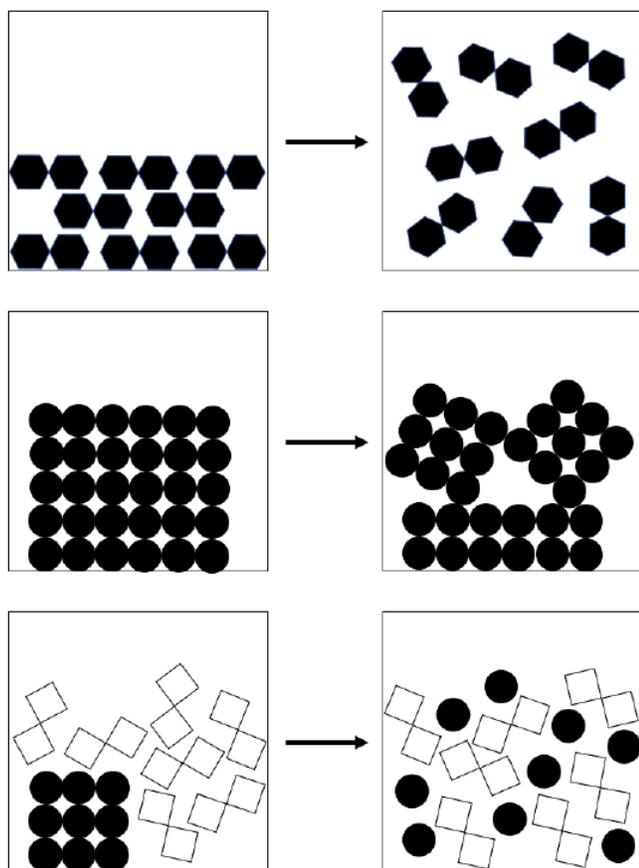
Figure 4.6.1: Nanoscopic Level Representations of Physical Changes

The first illustration represents a phase change. The left square represents a solid and the right square represents a gas. Since the same particle--consisting of two hexagons--makes up both the solid and the gas, this is a physical change. The second illustration represents a solid being broken into smaller pieces. The particle in both the left square and the right square consists of a single circle. Finally, the third illustration represents a solid dissolving in a liquid. Neither particle represented--the one formed from a single circle or the one formed from two squares--changes during the process of dissolving.

What all of these changes have in common is that the particles themselves remain unaltered though the transformation. Whatever particles you start with in the left square, you end with in the right square. Another way of saying this is that no new pure substances are formed.

Chemical Changes

Nanoscopic level representations of three chemical changes are shown below.
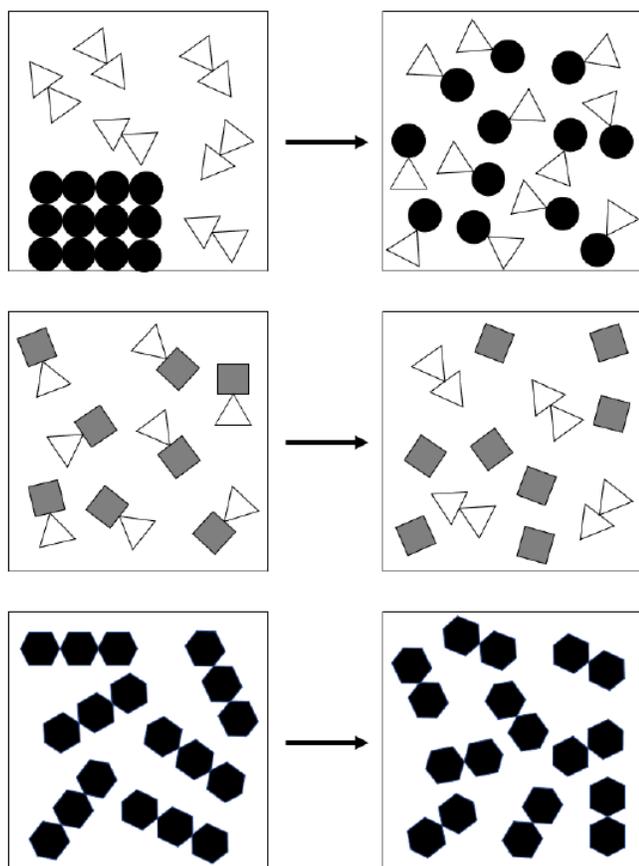
Figure 4.6.2: Nanoscopic Level Representations of Chemical Changes

What all these illustrations have in common is that new particles are formed while moving from the left square to the right square. In the first transformation a particle with a circle and a triangle is formed that did not originally exist. In the second illustration, a particle is formed that contains two triangles and another particle is formed that contains a single square. Originally, the only particle present contained a square and a triangle. Finally, a particle containing two hexagons is formed from a sample that originally contained particles with three hexagons.

The first illustration shows that a chemical change can sometimes include a change in state: here a solid disappears as it reacts and becomes part of the gas phase. The final illustration shows that it is possible for one sample of an element to change to another sample of an element. Even though the atoms themselves do not change, the way they bond together to form a particle is altered.

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4.7: Separating Mixtures through Physcial Changes

There are many techniques that use physical changes to separate mixtures. Here are a number of common separation techniques:

Distillation

Distillation is an effective method to separate mixtures comprised of two or more pure liquids. Distillation is a purification process where one component of a liquid mixture is vaporized and then condensed and isolated. In simple distillation Figure $4.7.1$, a mixture is heated and the most volatile component vaporizes at the lowest temperature. The vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state. The condensate that is collected is called distillate.

Outside the chemistry lab, distillation is commonly used to increase the alcohol content of liquors such as vodka, whiskey, and brandy.
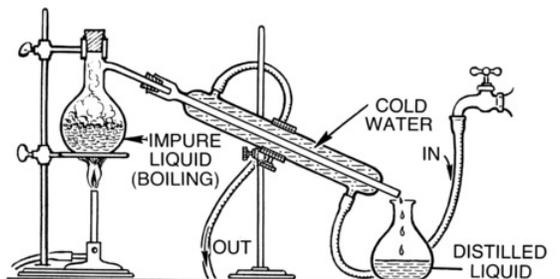
Figure $4.7.1$ *Distillation apparatus.*

Evaporation

Evaporation is a technique used to separate out a soluble solid from a liquid. The method drives off the liquid components from the solid components. The process typically involves heating the mixture until no more liquid remains. Evaporation can be used, for example, to obtain table salt from sea water, Figure $4.7.2$. The heat for the process comes from the sun.

Figure $4.7.2$ *Once the sea water in these evaporation ponds has evaporated, the salt can be harvested.*

Filtration

Filtration is a separation method used to separate out pure substances in mixtures comprised of particles, some of which are large enough in size to be captured with a porous material. Particle size can vary considerably, given the type of mixture. For instance, stream water is a mixture that contains naturally occurring biological organisms like bacteria, viruses, and protozoans. Some water filters can filter out bacteria, the length of which is on the order of 1 micron. Other mixtures, like soil, have relatively large particle sizes, which can be filtered through something like a coffee filter.
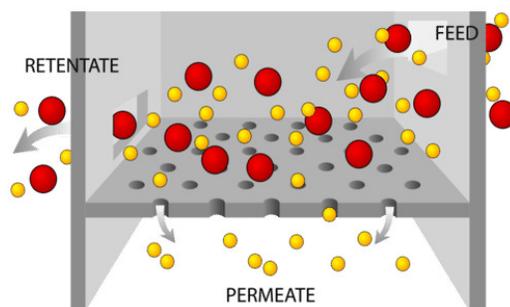
Figure $4.7.3$ *Filtration.*

Chromatography

Chromatography is the separation of a liquid or gas mixture by passing it through a medium in which the components move at different rates. Thin-layer chromatography is perhaps the simplist form of chromatography. It uses a plate--a plastic or glass slide covered on one side with silica (essentially very fine, purified sand), alumina, or some other solid. The plate is placed so that the bottom of it is submerged in a solvent, and the solvent then moves upward on the plate. If a drop of a mixture is placed a little above the level of the solvent, different components of the mixture will be carried upward at different rates depending on how strongly they interact with the silica. Figure $4.7.4$ shows separation of the chromophores (colored compounds) in a spinach leaf by thin layer chromatography.The pencil line at the bottom of the plate shows the original position of the mixture before the solvent started traveling up the plate.
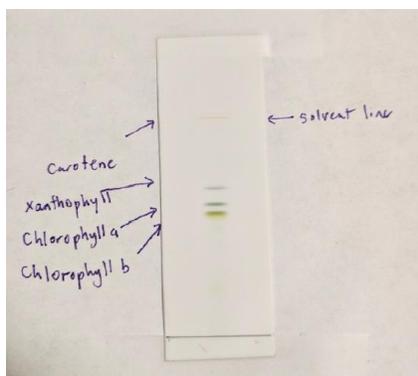
Figure $4.7.4$ Separation of compounds in a spinch leaf. (photo credit: Heather Coleman)

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4.8: Differences in Matter- Physical and Chemical Properties

> **◐ Learning Objectives**
>
> Given a property of matter, classify as physical or chemical.

All matter has physical and chemical properties. Physical properties are characteristics that scientists can measure without changing the composition of the sample under study, such as mass, color, and volume (the amount of space occupied by a sample). Chemical properties describe the characteristic ability of a substance to react to form new substances; they include its flammability and susceptibility to corrosion. All samples of a pure substance have the same chemical and physical properties. For example, pure copper is always a reddish-brown solid (a physical property) and always dissolves in dilute nitric acid to produce a blue solution and a brown gas (a chemical property).

Physical Property

A *physical property* is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets, a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. 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 4.8.1. 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 4.8.1: 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

Hardness helps determine how an element (especially a metal) might be used. Many elements are fairly soft (silver and gold, for example) while others (such as titanium, tungsten, and chromium) are much harder. Carbon is an interesting example of hardness. In graphite, (the "lead" found in pencils) the carbon is very soft, while the carbon in a diamond is roughly seven times as hard.

Figure 4.8.1: The graphite in a pencil (left) and the diamond in a diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

Melting and boiling points are somewhat unique identifiers, especially of compounds. In addition to giving some idea as to the identity of the compound, important information can be obtained about the purity of the material.

Chemical Properties

Chemical properties of matter describe its potential to undergo some chemical change or reaction by virtue of its composition. The elements, electrons, and bonds that are present give the matter potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, after studying chemistry for some time, you should be able to look at the formula of a compound and state some chemical property. For example, hydrogen has the potential to ignite and explode given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.
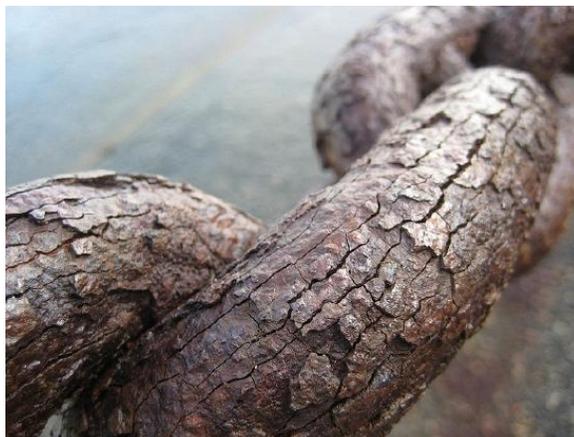
Figure 4.8.2: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure 4.8.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 4.8.2: Contrasting Physical and Chemical Properties

Physical Properties	Chemical Properties
Gallium metal melts at 30 ^{o}C.	Iron metal rusts.
Mercury is a very dense liquid.	A green banana turns yellow when it ripens.
Gold is shiny.	A dry piece of paper burns.

> ✔ **Example 4.8.1**
>
> Which of the following is a chemical property of iron?
>
>   a. Iron corrodes in moist air.
>   b. Density = 7.874 g/cm^3
>   c. Iron is soft when pure.
>   d. Iron melts at 1808 K.
>
> **Solution**
>
> "Iron corrodes in moist air" is the only chemical property of iron from the list.

> ❓ **Exercise 4.8.1A**
>
> Which of the following is a physical property of matter?
>
>   a. corrosiveness
>   b. pH (acidity)
>   c. density
>   d. flammability
>
> **Answer**
>
>   c

> ❓ **Exercise 4.8.1B**
>
> Which of the following is a chemical property?
>
>   a. flammability
>   b. melting point
>   c. boiling point
>   d. density
>
> **Answer**
>
>   a

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CHAPTER OVERVIEW

5: The Nuclei of Atoms

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5.1: The Periodic Table

> ### ◖ Learning Objectives
>
> - Given the name or symbol of an element and a periodic table, identify if the element is a:
>   - metal, nonmetal, or metalloid
>   - main group element or a transition metal
>   - alkali metal, alkaline earth metal, halogen, noble gas, or none of the above
> - Given the name or symbol of an element and a periodic table, identify the period and group number of the element.

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 5.1.1. 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.

Figure 5.1.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found here. (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).

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 5.1.2. 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 5.1.2, 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.
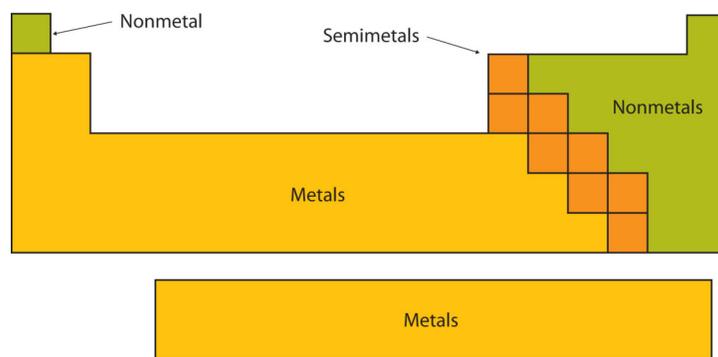
Figure 5.1.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

✔ **Example 5.1.1**

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 5.1.1, 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 5.1.1**

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 5.1.3. 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**.
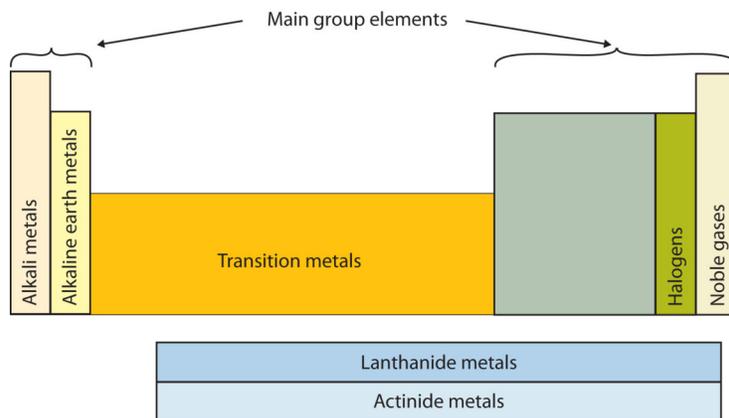

Figure 5.1.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

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 5.1.2: 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 5.1.2: 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 5.1.3: Classification of Elements
>
> Classify each element as metal, nonmetal, or metalloid.
>
>   a. Li
>   b. Ar
>   c. B
>   d. Fe

Solution

a. Lithium is a metal.
b. Argon is a nonmetal.
c. Boron is a metalloid.
d. Iron is a metal.

? Exercise 5.1.3: Classification of Elements

Classify each element as metal, nonmetal, or metalloid.

a. F
b. Si
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

Silicon is a metalloid.

Answer c:

Copper is a metal.

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5.2: Monotomic and Diatomic Elements

> **◑ Learning Objectives**
>
> - Name and write the symbols for the diatomic seven.

Molecular Elements

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called **molecules**. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

<div align="center">Table 5.2.1: Elements That Exist as Diatomic Molecules</div>

Hydrogen	Oxygen	Nitrogen	Fluorine	Chlorine	Bromine	Iodine

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules—a molecule with only two atoms (Table 5.2.1). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2, N_2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 (Figure 5.2.1). Notice that all the elements that occur as molecules are nonmetals.
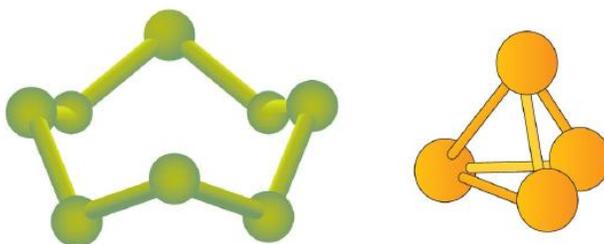

Figure 5.2.1: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 5.2.1 shows two examples of how molecules will be represented in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

Atomic Elements

While some elements exist as molecules, most elements exist with **individual atoms** as their basic unit. If an element is not one of those listed above as a molecular element, assume it is an atomic element. When writing the formula for an atomic element, only the symbol for the element appears (without any subscript).

> **✔ Example 5.2.1**
>
> Write the formula for a pure sample of each of the elements below.
>
> a. bromine
> b. neon
> c. hydrogen
> d. magnesium
> e. nitrogen

Solution

a. Br_2
b. Ne
c. H_2
d. Mg
e. N_2

States of the Elements

Gases

It is useful to know the states of the elements at the temperatures and pressures we are likely to encounter them (around $25^\circ C$ and average atmospheric pressure, also known as standard conditions). As the name of the group indicates, all of the noble gasses are gasses at these conditions. Additionally, fluorine, chlorine, oxygen, nitrogen, and hydrogen are gasses at standard conditions.

Liquids

Only two elements are liquids at standard conditions: bromine and mercury.

Solids

Most of the elements on the periodic table are solids at standard conditions. Unless an element was mentioned as being a gas or liquid above, assume it is a solid.

> ✔ Example 5.2.1
>
> Indicate the state of each of the elements below at $25^\circ C$ and normal atmospheric pressure.
>
> a. iodine
> b. carbon
> c. nitrogen
> d. bromine
> e. chlorine
>
> **Solution**
>
> a. solid
> b. solid
> c. gas
> d. liquid
> e. gas

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5.3: The Structure of Atoms

> **◑ Learning Objectives**
>
> - Name the three subatomic particles and provide their relative mass, charge, and location in the atom.

Atoms are not indivisible; they are composed of smaller components known as **subatomic particles**. Most chemistry can be explained by focusing on three subatomic particles: the electron, the proton, and the neutron.

Subatomic Particles

The **electron** is an extremely tiny particle with a mass of about 9.109×10^{-31} kg. It has a negative electrical charge. A **proton** has the same amount of charge as an electron, but its charge is positive. Another major difference between a proton and an electron is mass. Although still incredibly small, the mass of a proton is 1.673×10^{-27} kg, which is almost 2,000 times greater than the mass of an electron. Because opposite charges attract each other (while like charges repel each other), protons attract electrons (and vice versa). Finally, the **neutron** has about about the same mass as a proton but with no electrical charge; it is *neutral*. Table 5.3.1 lists some of the important characteristics of these subatomic particles and the symbols used to represent them.

Table 5.3.1: Properties of the Subatomic Particles

Particle	Symbol	Mass (kg)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1.673×10^{-27}	1	+1	nucleus
neutron	n^0	1.675×10^{-27}	1	0	nucleus
electron	e^-	9.109×10^{-31}	0.00055	−1	electron cloud

The Nucleus

Protons and neutrons are concentrated in a central region he called the **nucleus** (plural, *nuclei*) of the atom. The nucleus is incredibly small and dense. Since positively charged protons do not want to be directly next to each other (remember, like charges repel each other), a tremendous amount of energy binds the nucleus together.

Electrons are outside the nucleus in a region called the **electron cloud**. Within the electron cloud, electrons move around constantly. The areas occupied by various electrons and the relative energies associated with these areas is a major topic in General Chemistry. For now, it is adequate to think of electrons forming fuzzy clouds around nuclei (Figure 5.3.1). Electrons remain in the atom because they are attracted to the positive charge in the nucleus. Most of the mass of an atom is in the nucleus, while the electron cloud accounts for an atom's size. As a result, an atom consists largely of empty space.
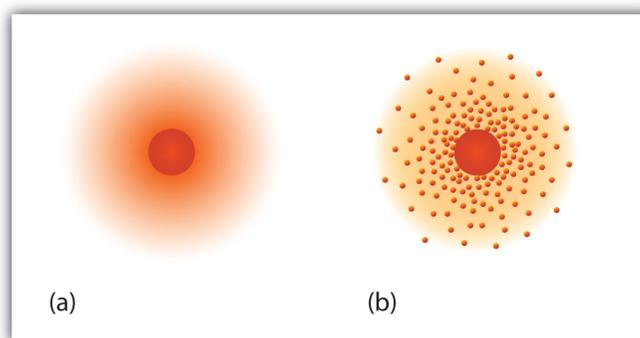

(a)   (b)

Figure 5.3.1: A Modern Depiction of Atomic Structure. A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. (a) The darker the color, the higher the probability that an electron will be at that point. (b) In a two-dimensional cross section of the electron in a hydrogen atom, the more crowded the dots, the higher the probability that an electron will be at that point. In both (a) and (b), the nucleus is in the center of the diagram.

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5.4: The Nuclei of Atoms

> ◉ **Learning Objectives**
>
> - Convert between a chemical species and the number of protons, neutrons, and electrons.
> - Given two or more chemical species determine if they are isotopes.

Now that we know how atoms are generally constructed, how many protons, neutrons, and electrons are in a specific kind of atom? First, if an atom is electrically neutral overall, then the number of protons equals the number of electrons. Because these particles have the same but opposite charges, equal numbers cancel out, producing a neutral atom.

Atomic Number

Experimentally, it is found that the magnitude of the positive charge in the nucleus of every atom of a particular element is the same. In other words, all atoms of the same element have the same number of protons. Furthermore, different elements have a different number of protons in their nuclei, so the number of protons in the nucleus of an atom is characteristic of a particular element. This discovery was so important to our understanding of atoms that the number of protons in the nucleus of an atom is called the **atomic number** (Z).

Based on its **atomic number**, you can determine the **number of protons in the nucleus** of an atom. For example, hydrogen has the atomic number 1; all hydrogen atoms have 1 proton in their nuclei. Helium has the atomic number 2; all helium atoms have 2 protons in their nuclei. There is no such thing as a hydrogen atom with 2 protons in its nucleus; a nucleus with 2 protons would be a helium atom. The atomic number *defines* an element. Take another look at the periodic table. Notice that for each element its atomic number is found above its symbol. Also notice that the periodic table is organized by increasing atomic number. The largest atoms have over 100 protons in their nuclei.

> ✔ **Example 5.4.1**
>
> What is the number of protons in the nucleus of each element?
>
> a. aluminum
> b. iron
> c. carbon
>
> **Answer a**
>
> According to Table 2.4.1, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.
>
> **Answer b**
>
> Iron has an atomic number of 26. Therefore, every iron atom has 26 protons in its nucleus.
>
> **Answer c**
>
> Carbon has an atomic number of 6. Therefore, every carbon atom has 6 protons in its nucleus.

> ❓ **Exercise 5.4.1**
>
> What is the number of protons in the nucleus of each element?
>
> a. sodium
> b. oxygen
> c. chlorine
>
> **Answer a**
>
> Sodium has 11 protons in its nucleus.
>
> **Answer b**

Oxygen has 8 protons in its nucleus.

Answer c

Chlorine has 17 protons in its nucleus

How many electrons are in an atom? Previously we said that for an electrically neutral atom, the **number of electrons equals the number of protons**, so the total opposite charges cancel. Thus, the **atomic number** of an element also gives the **number of electrons** in an atom of that element. (Later we will find that some elements may gain or lose electrons from their atoms, so those atoms will no longer be electrically neutral. Thus we will need a way to differentiate the number of electrons for those elements.)

> ✔ **Example 5.4.2**
>
> How many electrons are present in the atoms of each element?
>
> a. sulfur
> b. tungsten
> c. argon
>
> **Answer a**
>
> The atomic number of sulfur is 16. Therefore, in a neutral atom of sulfur, there are 16 electrons.
>
> **Answer b**
>
> The atomic number of tungsten is 74. Therefore, in a neutral atom of tungsten, there are 74 electrons.
>
> **Answer c**
>
> The atomic number of argon is 18. Therefore, in a neutral atom of argon, there are 18 electrons.

> ❓ **Exercise 5.4.2**
>
> How many electrons are present in the atoms of each element?
>
> a. magnesium
> b. potassium
> c. iodine
>
> **Answer a**
>
> Mg has 12 electrons.
>
> **Answer b**
>
> K has 19 electrons.
>
> **Answer c**
>
> I has 53 electrons.

Isotopes

How many neutrons are in atoms of a particular element? It turns out that atoms of the same element can have *different* numbers of neutrons. Atoms of the same element (i.e., same atomic number, Z) that have different numbers of neutrons are called isotopes. For example, 99% of the carbon atoms on Earth have 6 neutrons and 6 protons in their nuclei; about 1% of the carbon atoms have 7 neutrons in their nuclei. Naturally occurring carbon on Earth, therefore, is actually a mixture of isotopes, albeit a mixture that is 99% carbon with 6 neutrons in each nucleus.

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure 5.4.1 compares the three isotopes of hydrogen.
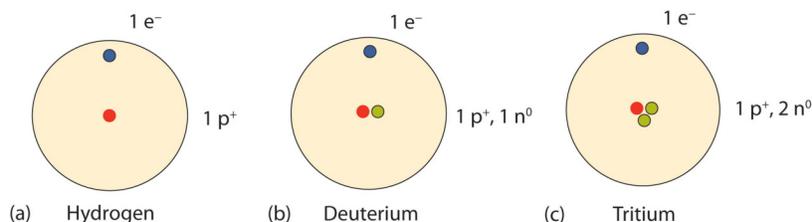
Figure 5.4.1: Isotopes of Hydrogen. Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

Most elements exist as mixtures of isotopes. In fact, there are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. The **mass number** (A) of an atom is the sum of the numbers of protons and neutrons in the nucleus. Given the mass number for a nucleus (and knowing the atomic number of that particular atom), you can determine the number of neutrons by subtracting the atomic number from the mass number.

A simple way of indicating the mass number of a particular isotope is to list it as a superscript on the left side of an element's symbol. Atomic numbers are often listed as a subscript on the left side of an element's symbol. Thus, we might see

$$\text{mass number} \longrightarrow {}^{56}_{26}\text{Fe} \atop \text{atomic number} \longrightarrow \qquad \qquad \tag{5.4.1}$$

which indicates a particular isotope of iron. The 26 is the atomic number (which is the same for all iron atoms), while the 56 is the mass number of the isotope. To determine the number of neutrons in this isotope, we subtract 26 from 56: 56 − 26 = 30, so there are 30 neutrons in this atom.

✔ Example 5.4.3

How many protons and neutrons are in each atom?

a. ${}^{35}_{17}\text{Cl}$
b. ${}^{127}_{53}\text{I}$

Answer a

In ${}^{35}_{17}\text{Cl}$ there are 17 protons, and 35 − 17 = 18 neutrons in each nucleus.

Answer b

In ${}^{127}_{53}\text{I}$ there are 53 protons, and 127 − 53 = 74 neutrons in each nucleus.

? Exercise 5.4.3

How many protons and neutrons are in each atom?

a. ${}^{197}_{79}\text{Au}$
b. ${}^{23}_{11}\text{Na}$

Answer a

In ${}^{197}_{79}\text{Au}$ there are 79 protons, and 197 − 79 = 118 neutrons in each nucleus.

Answer b

In ${}^{23}_{11}\text{Na}$ there are 11 protons, and 23 − 11 = 12 neutrons in each nucleus.

It is not absolutely necessary to indicate the atomic number as a subscript because each element has its own unique atomic number. Many isotopes are indicated with a superscript only, such as ${}^{13}\text{C}$ or ${}^{235}\text{U}$. You may also see isotopes represented in print as, for

example, carbon-13 or uranium-235.

5.5: Atomic Masses

> ### Learning Objectives
>
> - Given the isotopic abundance and atomic mass of each isotope of an element, determine the average atomic mass.

Even though atoms are very tiny pieces of matter, they have mass. Their masses are so small, however, that chemists often use a unit other than grams to express them—the atomic mass unit.

Atomic Mass Unit

The atomic mass unit (abbreviated u, although amu is also used) is defined as 1/12 of the mass of a ^{12}C atom:

$$1\,\text{u} = \frac{1}{12} \text{ the mass of } ^{12}\text{C atom} \tag{5.5.1}$$

It is equal to 1.661×10^{-24} g.

Masses of other atoms are expressed with respect to the atomic mass unit. For example, the mass of an atom of ^{1}H is 1.008 u, the mass of an atom of ^{16}O is 15.995 u, and the mass of an atom of ^{32}S is 31.97 u. Note, however, that these masses are for particular isotopes of each element. Because most elements exist in nature as a mixture of isotopes, any sample of an element will actually be a mixture of atoms having slightly different masses (because neutrons have a significant effect on an atom's mass). How, then, do we describe the mass of a given element? By calculating an average of an element's atomic masses, weighted by the natural abundance of each isotope, we obtain a weighted average mass called the atomic mass (also commonly referred to as the *atomic weight*) of an element.

Atomic Mass is the Weighted Average Mass of Isotopes

As stated above, most elements occur naturally as a mixture of two or more isotopes. Listed below (Table 5.5.1) are the naturally occurring isotopes of selected elements along with the percent natural abundance of each.

Table 5.5.1: Atomic Masses and Percent Abundances of Some Natural Isotopes

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Hydrogen	$^{1}_{1}$H	99.985	1.0078	1.0079
	$^{2}_{1}$H	0.015	2.0141	
	$^{3}_{1}$H	negligible	3.0160	
Carbon	$^{12}_{6}$C	98.89	12.000	12.011
	$^{13}_{6}$C	1.11	13.003	
	$^{14}_{6}$C	trace	14.003	
Oxygen	$^{16}_{8}$O	99.759	15.995	15.999
	$^{17}_{8}$O	0.037	16.995	
	$^{18}_{8}$O	0.204	17.999	
Chlorine	$^{35}_{17}$Cl	75.77	34.969	35.453
	$^{38}_{17}$Cl	24.23	36.966	
Copper	$^{63}_{29}$Cu	69.17	62.930	63.546
	$^{65}_{29}$Cu	30.83	64.928	

For some elements, one particular isotope is much more abundant than any other isotopes. For example, 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 substantial quantities. Chlorine (atomic number 17) is 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 approximately 36.

$$\frac{34.969\,u + 36.966\,u}{2} = 35.968\,u \tag{5.5.2}$$

As you can see, the average atomic mass given in the last column of the table above (35.453) is significantly lower. Why? The reason is that we need to take into account the natural abundance percentages of each isotope in order to calculate what is called 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.*

$$0.7577\,(34.969\,u) + 0.2423\,(36.966\,u) = 35.453\,u \tag{5.5.3}$$

The weighted average is determined by multiplying the percent of natural abundance by the actual mass of the isotope. This is repeated until there is a term for each isotope. For chlorine, there are only two naturally occurring isotopes so there are only two terms.

Atomic mass = (%1)(mass 1) + (%2)(mass 2) + ⋯

Another example: oxygen exists as a mixture that is 99.759% ^{16}O, 0.037% ^{17}O and 0.204% ^{18}O. The atomic mass of oxygen (use percent natural abundance data from Table 2.5.1) would be calculated as follows:

Atomic mass = (%1)(mass 1) + (%2)(mass 2) + (%3)(mass 3)

$$0.99759\,(15.995u) + 0.00037\,(16.995u) + 0.00204\,(17.999u) = 15.999u \tag{5.5.4}$$

To confirm your answer, compare the calculated value to the weighted mass displayed on the periodic table.

> ✔ **Example 5.5.1**
>
> Calculate the atomic mass of oxygen. Oxygen exists as a mixture of 3 isotopes. Their respective masses and natural abundance are shown below.
>
> - ^{16}O: 15.995 u (99.759%)
> - ^{17}O: 16.995 u (0.037%)
> - ^{18}O: 17.999 u (0.204%)
>
> **Solution**
>
> Multiply the isotope abundance by the actual mass of the isotope, and then sum up the products.
>
> $$0.99759\,(15.995\,u) + 0.00037\,(16.995\,u) + 0.00204\,(17.999\,u) = 15.999\,u$$

> **Exercise 5.5.1**
>
> Calculate the atomic mass of copper. Copper exists as a mixture of 2 isotopes. Their respective masses and natural abundance are shown below.
>
> - ^{63}Cu: 62.930 u (69.17%)
> - ^{65}Cu: 64.928 u (30.83%)
>
> **Answer**
>
> 63.546 u

The **atomic mass** of each element is found under the element symbol in most **periodic tables.** Examples are shown below. The atomic mass of tin (Sn) is 118.71 u while the atomic mass of carbon (C) is 12.011 u. On the other hand, the **atomic number (Z)** of each element is found **above** the atomic symbol.
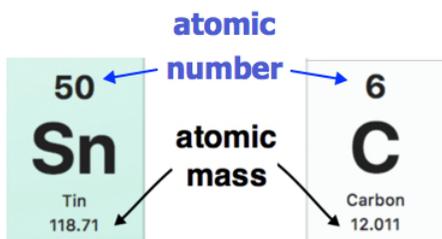
Atomic mass indicated on entries of the Periodic Table. (public Domain; Pubchem)

The **periodic table** is found in this link:

https://pubchem.ncbi.nlm.nih.gov/periodic-table/png/Periodic_Table_of_Elements_w_Atomic_Mass_PubChem.png

> ✔ Example 5.5.2: Mass of Carbon
>
> What is the average mass of a carbon atom in grams? The atomic mass is found in the Periodic Table. Please use two decimal places.
>
> **Solution**
>
> This is a simple one-step conversion, similar to conversions we did in Chapter 1. We use the fact that 1 u = 1.661×10^{-24} g:
>
> $$12.01 \; \cancel{u} \times \frac{1.661 \times 10^{-24} \text{ g}}{1 \; \cancel{u}} = 1.995 \times 10^{-23} \text{ g}$$
>
> This is an extremely small mass, which illustrates just how small individual atoms are.

> ? Exercise 5.5.2: Mass of Tin
>
> What is the average mass of a tin atom in grams? The average atomic mass is found in the Periodic Table. Please use two decimal places.
>
> **Answer**
>
> $$118.71 \; \cancel{u} \times \frac{1.661 \times 10^{-24} \text{ g}}{1 \; \cancel{u}} = 1.972 \times 10^{-22} \text{ g}$$

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5.6: Bohr Diagrams of Atoms and Ions

> **◖◗ Learning Objectives**
>
> - Draw the Bohr Diagram of an atom with 18 electrons or fewer.

Electron Shells

Niels Bohr proposed an early model of the atom as a central nucleus containing protons and neutrons being orbited by electrons in shells. As previously discussed, there is a connection between the number of protons in an element, the atomic number that distinguishes one element from another, and the number of electrons it has. In all electrically-neutral atoms, the number of electrons is the same as the number of protons. Each element, when electrically neutral, has a number of electrons equal to its atomic number.

An early model of the atom was developed in 1913 by Danish scientist Niels Bohr (1885–1962). The Bohr model shows the atom as a central nucleus containing protons and neutrons with the electrons in circular orbitals at specific distances from the nucleus (Figure 5.6.1). These orbits form electron shells or energy levels, which are a way of visualizing the number of electrons in the various shells. These energy levels are designated by a number and the symbol "n." For example, the 1n shell represents the first energy level located closest to the nucleus.
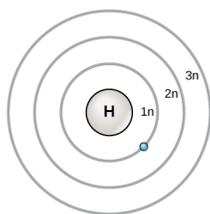
Figure 5.6.1: The Bohr model postulated that electron orbited the nucleus in shells of fixed distance.

An electron normally exists in the lowest energy shell available, which is the one closest to the nucleus. Energy from a photon of light can bump it up to a higher energy shell, but this situation is unstable and the electron quickly decays back to the ground state.

Bohr Diagrams

Bohr diagrams show electrons orbiting the nucleus of an atom somewhat like planets orbit around the sun. In the Bohr model, electrons are pictured as traveling in circles at different shells, depending on which element you have. Figure 5.6.2 contrast the Bohr diagrams for lithium, fluorine and aluminum atoms. The shell closest to the nucleus is called the K shell, next is the L shell, next is the M shell.
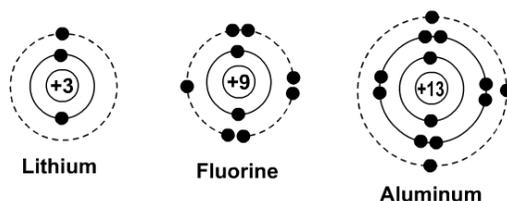
Figure 5.6.2: Bohr diagrams for neutral lithium, fluorine and aluminum atoms.

Each shell can only hold certain number of electrons. K shell can have 2, L can have 8 , M can have 18 electrons and so on.

- Lithium has three electrons:
  - two go to K shell and
  - the remaining one goes to the L shell.
  - Its electronic configuration is K(2), L(1)

- Fluorine has nine electrons:
  - two go to K shell and
  - the remaining seven go to the L shell.
  - Its electronic configuration is K(2), L(7). Note that L can have 8 electrons.

- Aluminum has thirteen electrons:
  - two go to the K shell,
  - eight go to the L shell, and
  - remaining three go to the M shell.
  - Its electronic configuration is K(2), L(8), M(3). Note that the M shell can have 18 electrons.

Orbitals in the Bohr model

Electrons fill orbit shells in a consistent order. Under standard conditions, atoms fill the inner shells (closer to the nucleus) first, often resulting in a variable number of electrons in the outermost shell. The innermost shell has a maximum of two electrons, but the next two electron shells can each have a maximum of eight electrons. This is known as the octet rule which states that, with the exception of the innermost shell, atoms are more stable energetically when they have eight electrons in their valence shell, the outermost electron shell. Examples of some neutral atoms and their electron configurations are shown in Figure 5.6.3. As shown, helium has a complete outer electron shell, with two electrons filling its first and only shell. Similarly, neon has a complete outer 2n shell containing eight electrons. In contrast, chlorine and sodium have seven and one electrons in their outer shells, respectively. Theoretically, they would be more energetically stable if they followed the octet rule and had eight.
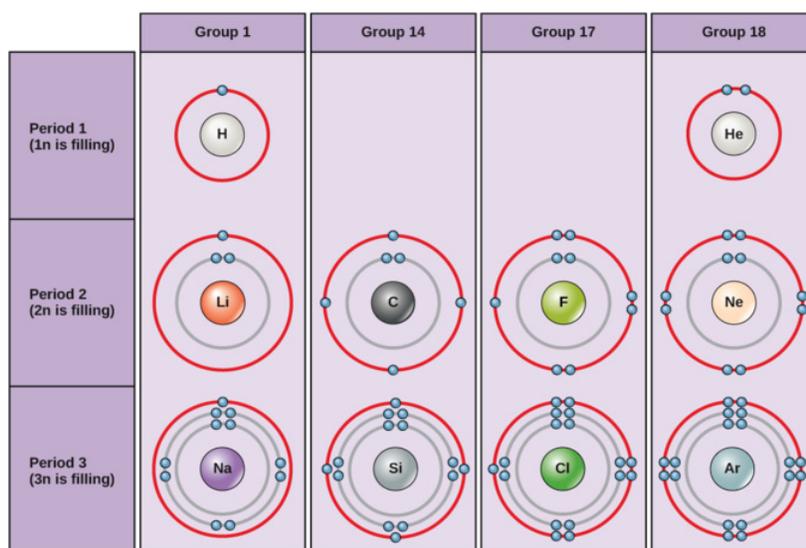
Figure 5.6.3:

Bohr diagrams

Bohr diagrams indicate how many electrons fill each principal shell. Group 18 elements (helium, neon, and argon are shown) have a full outer, or valence, shell. A full valence shell is the most stable electron configuration. Elements in other groups have partially-filled valence shells and gain or lose electrons to achieve a stable electron configuration.

An atom may gain or lose electrons to achieve a full valence shell, the most stable electron configuration. The periodic table is arranged in columns and rows based on the number of electrons and where these electrons are located, providing a tool to understand how electrons are distributed in the outer shell of an atom. As shown in , the group 18 atoms helium (He), neon (Ne), and argon (Ar) all have filled outer electron shells, making it unnecessary for them to gain or lose electrons to attain stability; they are highly stable as single atoms. Their non-reactivity has resulted in their being named the inert gases (or noble gases). In comparison, the group 1 elements, including hydrogen (H), lithium (Li), and sodium (Na), all have one electron in their outermost shells. This means that they can achieve a stable configuration and a filled outer shell by donating or losing an electron. As a result of losing a negatively-charged electron, they become positively-charged ions. When an atom loses an electron to become a positively-charged ion, this is indicated by a plus sign after the element symbol; for example, Na^+. Group 17 elements, including fluorine and chlorine, have seven electrons in their outermost shells; they tend to fill this shell by gaining an electron from other atoms, making them negatively-charged ions. When an atom gains an electron to become a negatively-charged ion this is indicated by a minus sign after the element symbol; for example, F^-. Thus, the columns of the periodic table represent the potential shared state of these elements' outer electron shells that is responsible for their similar chemical characteristics.

5.7: Periodic Trends- Atomic Radius

> ◀▶ **Learning Objectives**
>
> - Using a Periodic Table, state the trend in atomic size within a group or period of elements.

Many events draw large numbers of people to them. Even an outdoor event can fill up so that there is not room for anyone else. 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. More people can fit into a given space than can elephants, because elephants are larger than people. More squirrels can fit into that same space than can people for the same reason. Knowing the sizes of objects to be dealt with is important when 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 scientists to 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.
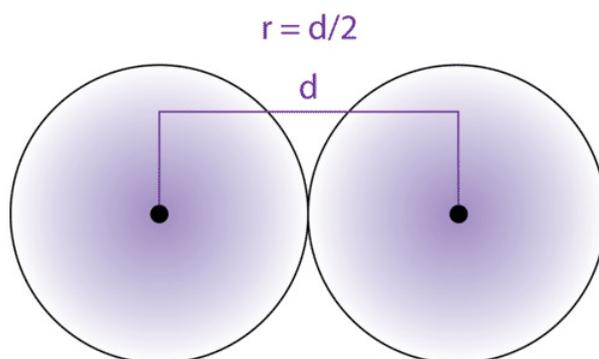
Figure 5.7.1: The atomic radius (r) of an atom can be defined as one half the distance (d) between two nuclei in a diatomic molecule.

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 $\frac{74}{2} = 37\text{ pm}$.
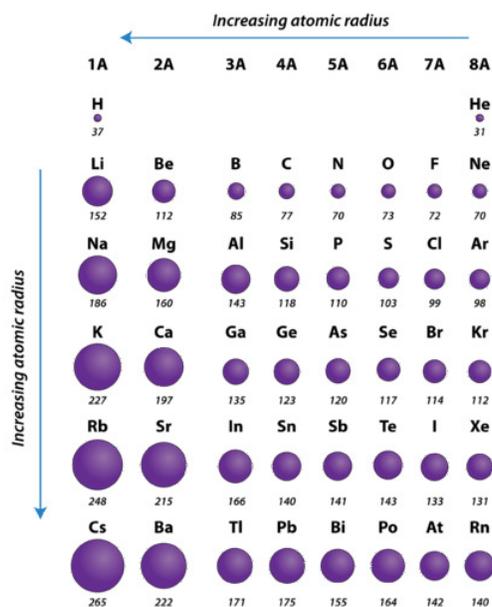
Figure 5.7.2: Atomic radii of the representative elements measured in picometers.

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 5.7.3: 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.

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5.8: Periodic Trends- Ionization Energy

> **◖◗ Learning Objectives**
>
> - Students will be able to describe the term ionization energy on a quiz.
> - The student will be able to state the trend in ionization energy within a group or period of elements.

Like many other animals, sheep travel in herds. The tendency is for each individual sheep to stay with the herd. However, a sheep may sometimes wander off, depending on how strong the attraction is for a particular food or water supply. At other times, a sheep may become frightened and run off. Whether a sheep chooses to stay with the herd or go its own way depends on the balance between attraction to the herd and attraction to an outside influence.

There is an ongoing tension between the electrons and protons in an atom. Reactivity of the atom depends in part on how easily the electrons can be removed from the atom. We can measure this quantity and use it to make predictions about the behavior of atoms.

Ionization Energy

Ionization energy is the energy required to remove an electron from a specific atom. It is measured in kJ/mol, which is an energy unit, much like calories. The ionization energies associated with some elements are described in the table below. For any given atom, the outermost valence electrons will have lower ionization energies than the inner-shell kernel electrons. As more electrons are added to a nucleus, the outer electrons become shielded from the nucleus by the inner shell electrons. This is called **electron shielding**.

Table 5.8.1: Ionization Energies (kJ/mol) of the First 8 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
H	1312					
He	2373	5251				
Li	520	7300	11,815			
Be	899	1757	14,850	21,005		
B	801	2430	3660	25,000	32,820	
C	1086	2350	4620	6220	38,000	47,261
N	1400	2860	4580	7500	9400	53,000
O	1314	3390	5300	7470	11,000	13,000

The following trend is seen when we plot the first ionization energies vs. atomic number for the main group elements:
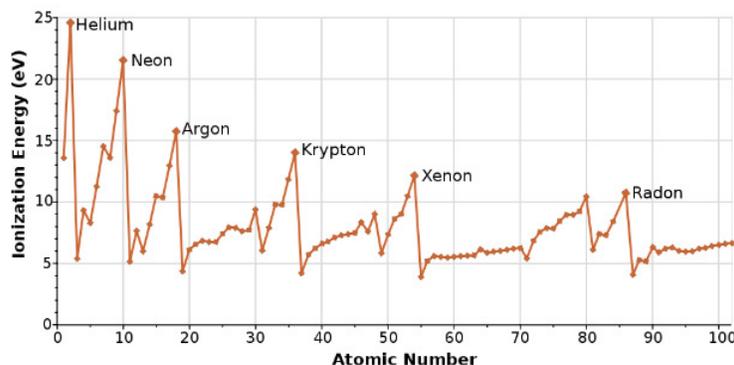
Figure 5.8.1: Ionization energy and atomic number.

Moving from left to right across the periodic table, the ionization energy for an atom increases. We can explain this by considering the nuclear charge of the atom. The more protons in the nucleus, the stronger the attraction of the nucleus to electrons. This

stronger attraction makes it more difficult to remove electrons.

Within a group, the ionization energy decreases as the size of the atom gets larger. On the graph, we see that the ionization energy increases as we go up the group to smaller atoms. In this situation, the first electron removed is farther from the nucleus as the atomic number (number of protons) increases. Being farther away from the positive attraction makes it easier for that electron to be pulled off.
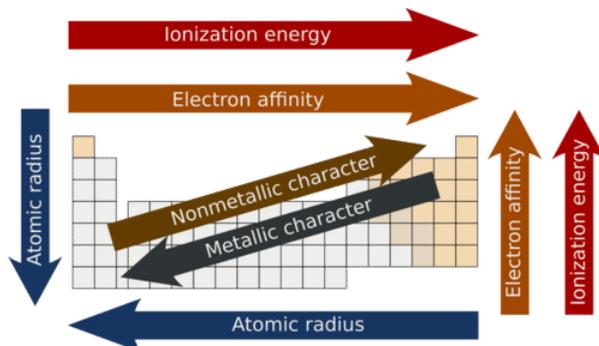
Figure 5.8.2: Periodic trends.

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5.9: Radioactivity

> **◑ Learning Objectives**
>
> - Identify to symbols for alpha, beta and gamma radiation.
> - List the order if increasing penetrating power: alpha, beta and gamma radiation.

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:

$$_{6}^{12}\text{C} \tag{5.9.1}$$

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}H and ^{3}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 $_2^4$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:

$$_{92}^{235}\text{U} \rightarrow \, _2^4\text{He} + \, _{90}^{231}\text{Th} \tag{5.9.2}$$

How do we know that a product of the reaction is $_{90}^{231}$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, $_{90}^{231}$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, $_{92}^{235}$U is the parent isotope, and $_{90}^{231}$Th is the daughter isotope. When one element changes into another in this manner, it undergoes *radioactive decay*.

> **✔ Example 5.9.1: Radon-222**
>
> 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 $_{86}^{222}$Rn. We represent the alpha particle as $_2^4$He and use subtraction (222 − 4 = 218 and 86 − 2 = 84) to identify the daughter isotope as an isotope of polonium, $_{84}^{218}$Po:
>
> $$_{86}^{222}\text{Rn} \rightarrow \, _2^4\text{He} + \, _{84}^{218}\text{Po}$$
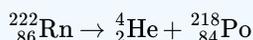
> **? Exercise 5.9.1: Polonium-209**
>
> Write the nuclear equation that represents the radioactive decay of polonium-209 by alpha particle emission and identify the daughter isotope.
>
> **Answer**
>
> $$^{209}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{205}_{82}\text{Pb}$$
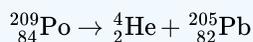
The second major type of radioactive emission is called a beta particle, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1− charge. We can also represent a beta particle as $_{-1}^{0}\text{e}$ or 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:

$$^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e} \tag{5.9.3}$$

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:

$$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He} + \gamma \tag{5.9.4}$$

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 5.9.2: Boron-12**
>
> 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, $_{-1}^{0}\text{e}$. 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:
>
> $$^{12}_{5}\text{B} \rightarrow {}^{12}_{6}\text{C} + {}^{0}_{-1}\text{e} + \gamma \tag{5.9.5}$$
>
> The daughter isotope is $^{12}_{6}\text{C}$.

> **? Exercise 5.9.2: Iodine-131**
>
> 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**
>
> $$^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}\text{e} + \gamma \tag{5.9.6}$$

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 5.9.1 summarizes the properties of the three main types of radioactive emissions.
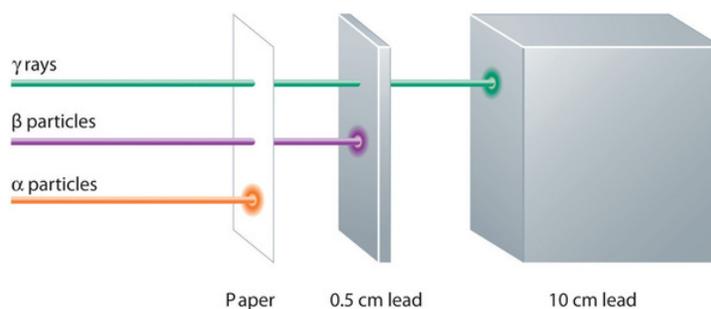
Figure 5.9.2: Different emissions exhibit different pentration powers. (CC BY-NC-SA 3.0; anonymous)

Table 5.9.1: The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, ^4_2He	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

$$^{235}_{92}\text{U} \rightarrow \, ^{139}_{56}\text{Ba} + \, ^{94}_{36}\text{Kr} + 2\,^1_0\text{n} \tag{5.9.7}$$

where ^1_0n 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.

CHAPTER OVERVIEW

6: Ions, Ionic Bonding, and the Nomenclature of Ionic Compounds

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6.1: Ions

> 🧠 **Learning Objectives**
>
> - Predict the charge of monatomic main group ions.
> - Given the symbol for a monatomic, fixed-charge ion write the systematic name.
> - Given the systematic name for a monatomic, fixed-charge ion write the symbol.
> - Move between a symbol or name for a monatomic, fixed-charge ion and the number of electrons.

Cations and Anions

Recall that neutral atoms have the same number of protons and electrons. In nature, however, many atoms lose or gain electrons and change from being neutral to having a net positive or negative charge. An atom that picks up a charge by gaining or losing electrons is called an **ion**. In this section we will focus on **monatomic ions**: ions formed from a single atom. Later in this chapter we will look at a second type of ion called a **polyatomic ion**; these form when a group of bonded atoms collectively gain a charge.

Metals have a strong tendency to lose electrons. When a metal atom loses electrons, it ends up with a greater number of protons than electrons. The atom therefore now has a positive charge because of the overabundance of protons. Any ion with a positive charge is referred to as a **cation**. Nonmetals, on the other hand, tend to gain electrons. This means the ion formed from a nonmetal atom with end up with more electrons than protons and have an overall negative charge. An **anion** is the general name given to any negatively charged ion. Hydrogen behaves a little differently than the other nonemtals. Most often it loses its only electron and becomes a cation. There are some rare cases, however, where it can gain an electron an become an anion.

> The names for positive and negative ions are pronounced CAT-eye-ons (cations) and ANN-eye-ons (anions), respectively.

Formation of Ions

For any ion, we can find the charge by taking the number of electrons and subtracting it from the number of protons. We will first look at a sodium atom and a sodium ion to understand this relationship.

Sodium atoms have 11 protons (because the atomic number of sodium is 11) and 11 electrons (because the number of electrons must equal the number of protons in a neutral atom). Subtracting 11 (the number of electrons) from 11 (the number of protons) gives zero, the overall charge on the neutral atom. See the right side of Figure 6.1.1.

Sodium atoms always lose one electron when they become ions. This means they still have 11 protons, but now have only 10 electrons. Since 11 minus 10 is one, the overall charge on the ion is +1. You always need to indicate the sign (+ or -) when writing charges. Additionally, to differentiate a sodium atom from a sodium ion, both the chemical formula and the name change. The name change is rather trivial for cations: you just add the word *ion* after the name of the element. For the chemical formula of an ion, the charge is indicated above and to the right of the symbol for the element. The chemical formula for a sodium ion is therefore Na^+. When the charge is plus one, the one is left out of the chemical formula. To give another example, calcium atoms always lose two electrons when they form ions. The name of the resulting ion is a calcium ion and its chemical formula is Ca^{2+}.
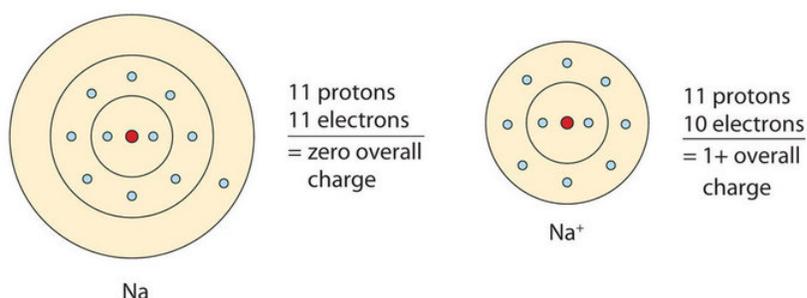
Figure 6.1.1: The Formation of a sodium ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a +1 charge.

Now, let's consider a chlorine atom. A chlorine atom always gains one electron when it forms an ion (Figure 6.1.2). A chlorine atom starts with 17 electrons and 17 protons and is neutral. After gaining an electron to become an ion, it now has 18 electrons. The

charge is determined by taking 17 (the number of protons) and subtracting 18 (the number of electrons); it is equal to -1. We have to include the sign to indicate that it is negative.

The chemical formula for a chlorine atom is Cl and the chemical formula for the ion is Cl⁻. Notice again how the one is left off of the ion charge when writing the formula. Anion names work slightly differently than cation names: the ion formed from a chlorine atom is called a chloride ion. For monatomic anions, the suffix *-ide* is added onto the root name of the element to create the name of the anion. Oxygen atoms, for example, become oxide ions (O^{2-}) and phosphorus atoms become phosphide ions (P^{3-}).
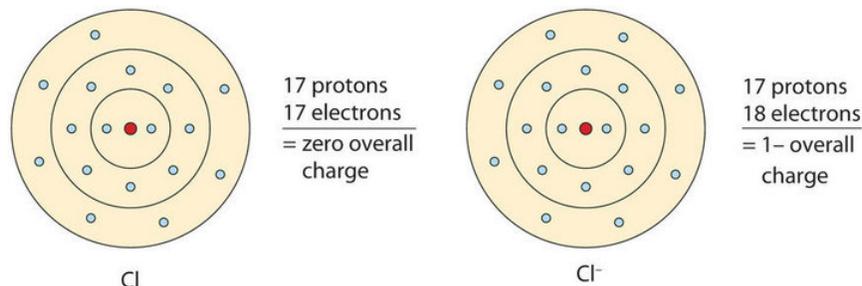
Figure 6.1.2: The Formation of a chlorine ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a -1 charge.

✔ **Example 6.1.1**

For each ion, determine its name and the number of electrons it contains.

 a. Mg^{2+}
 b. N^{3-}
 c. F^-

Solution

(a) Since Mg^{2+} is a cation, its name is the name of the element it comes from plus the word *ion*. This would make it a magnesium ion. Magnesium atoms contain 12 protons in their nucleus. To get a +2 charge, the ion had to lose two electrons. A magnesium ion therefore has 10 electrons.

(b) Since N^{3-} is an anion, its name is the root name of the element with the suffix *-ide*. This would make it a nitride ion. Nitrogen atoms contain 7 protons in their nucleus. To get a -3 charge, the ion had to gain three electrons. A nitride ion therefore has 10 electrons.

(b) Since F^- is an anion, its name is the root name of the element with the suffix *-ide*. This would make it a fluoride ion. Fluorine atoms contain 9 protons in their nucleus. To get a -1 charge, the ion had to gain one electron. A fluoride ion therefore has 10 electrons.

Exercise 6.1.1

For each ion, determine its name and the number of electrons it contains.

 a. S^{2-}
 b. Al^{3+}
 c. Rb^+

Answer

> (a) sulfide ion; 18 electrons

> (b) aluminum ion, 10 electrons

> (c) rubidium ion, 36 electrons

Charges of Main Group Ions

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge. 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 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}.

Notice that the transition metals and most metals in groups 13 through 16 (3A through 6A) do not appear in Figure 6.1.3. This is because the charges on these metal ions do not follow such a regular pattern or because it is possible for the metals to form multiple ions with different charges. These cations will be discussed in the next section.

Figure 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 6.1.2

For each element, write the chemical formula and the name of the ion it forms.

a. iodine
b. potassium
c. selenium

Solution

(a) I^-, iodide ion

(b) K^+, potassium ion

(c) Se^{2-}, selenide ion

Exercise 6.1.2

For each element, write the chemical formula and the name of the ion it forms.

a. strontium
b. bromine
c. lithium

Answer

(a) Sr^{2+}, strontium ion

(b) Br^-, bromide ion

(c) Li^+, lithium ion

6.2: Ions With Variable Charges

> **◐ Learning Objectives**
>
> - Given the symbol for a monatomic, variable-charge ion write the systematic name.
> - Given the systematic name for a monatomic, variable-charge ion write the symbol.

Up until now, we have looked at elements that form ions with fixed charges. By fixed charge, we mean that we can always predict the charge on the ion. An oxide ion, for example, always has a charge of -2. It is impossible for oxygen to form an ion with a -1 charge or a -3 charge.

Many metals, on the other hand, form ions that have variable charges. Saying an ion has variable charges means that the same element can form ions with different charges. For example, it is possible for copper to form both an ion with a +1 charge and to form an ion with a +2 charge. The large majority of transition metals form ions with variable charges. There is also a small block of main group metals that form ions with variable charges as well (see Figure 6.2.1 below). There is no need to memorize the charges of metals that form ions with variable charges, as the charges may be identified by using Roman numerals. When writing the name of the ion, the name of the metal is immediately followed by parentheses (no space) within which is a Roman numeral equal to the magnitude of charge on the ion.

For example,

- the iron(II) ion has a chemical formula of Fe^{2+}.
- the iron(III) ion has a chemical formula of Fe^{3+}.

Figure 6.2.1: The periodic table showing elements that form ions with variable charges.

For a reminder about how to write Roman numerals, see Table 6.2.1 below. Be especially careful with four (IV in Roman numerals) and six (VI in Roman numerals).

charge	Roman numeral
+1	I
+2	II
+3	III
+4	IV
+5	V
+6	VI

Table 6.2.1: Roman numerals used to represent charges.

? ✏️ Exercise 6.2.1

If the chemical formula of an ion is provided, write the name. If the name of an ion is provided, write the chemical formula.

A. Mn^{3+}
B. copper(I) ion
C. lead(IV) ion
D. Cu^{2+}

Answer A

manganese(III) ion

Answer B

Cu^+

Answer C

Pb^{4+}

Answer D

copper(II) ion

When writing names for compounds containing metals that form ions with variable charges, the name *must include* the Roman numeral that shows the charge on the metal ion. It is incorrect to use the term *chromium ion*, for example, since chromium can form both an ion with a +2 charge and an ion with a +3 charge. Roman numerals must *not* be included when the metal only forms ions with fixed charges. It is not correct, for example, to call a calcium ion a calcium(II) ion. Table 6.2.2 lists the cations that form ions with fixed charges. These need to be memorized.

ion	charge
any group 1 ion	+1
silver ion	+1
any group 2 ion	+2
cadmium ion	+2
zinc ion	+2
aluminum ion	+3
gallium ion	+3
scandium ion	+3

Table 6.2.2: Ions with Fixed Charges

? ✏️ Exercise 6.2.2

If the formula of an ion is provided, write the name. If the name of an ion is provided, write the formula.

A. Zn^{2+}
B. Co^{3+}
C. Sn^{2+}
D. Ag^+

Answer A

zinc ion

Answer B

cobalt(III) ion

Answer C

tin(II) ion

Answer D

silver ion

Contributions & Attributions

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6.3: Polyatomic Ions

> **◖❙ Learning Objectives**
>
> - Given the symbol for a polyatomic ion write the systematic name.
> - Given the systematic name for a polyatomic ion write the symbol.

Polyatomic Ions

Some ions consist of groups of atoms *covalently* bonded together and have an **overall electric charge**. Covalent bonding will be discussed in the next chapter. Because these ions contain more than one atom, they are called polyatomic ions. The structures, names and formulas of some polyatomic ions are found in Table 3.3.1.

Table 6.3.1: *Some Polyatomic Ions*

Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table 6.3.2 lists 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 6.3.2 lists the most common polyatomic ions.

Table 6.3.2: Ion Names and Ion Formulas of Common Polyatomic Ions

Ion Name	Ion Formula
ammonium ion	NH_4^+
hydroxide ion	OH^-
cyanide ion	CN^-

Ion Name	Ion Formula
carbonate ion	CO_3^{-2}
bicarbonate or hydrogen carbonate	HCO_3^-
acetate ion	$C_2H_3O_2^-$ **or** $CH_3CO_2^-$
chlorate ion	ClO_3^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
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^+), 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. 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, 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.

6.4: The Crystaline Struture of Ionic Compounds

> **① Learning Objectives**
>
> - Describe ionic compounds as an extended three dimensional array (a lattice structure).
> - Explain that the formula of an ionic compound specifies the ratio of cations to anions.
> - Given the name of an ionic compound (which may contain a fixed-charge or variable charge metal and may contain a polyatomic ion) write the formula.

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 2 metal) and a nonmetallic element (bromine, a group 17 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 15, and oxygen, from group 16). It is not an ionic compound; it belongs to the category of covalent compounds discussed 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 $Ba(NO_3)_2$, you may recognize the "NO₃" part as the nitrate ion, 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 Ba^{2+} ion, with the +2 charge balancing the overall −2 charge from the two nitrate ions. Thus, this compound is also ionic.

> **✔ Example 6.4.3**
>
> Identify each compound as ionic or not ionic.
>
> a. Na_2O
> b. PCl_3
> c. NH_4Cl
> d. OF_2
>
> **Solution**
>
> a. Sodium is a metal, and oxygen is a nonmetal; therefore, Na_2O is expected to be ionic.
> b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.
> c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic.
> d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.

> **? Exercise 6.4.3**
>
> Identify each compound as ionic or not ionic.
>
> a. N_2O
> b. $FeCl_3$
> c. $(NH_4)_3PO_4$
> d. $SOCl_2$
>
> **Answer a:**
>
> not ionic
>
> **Answer b:**
>
> ionic
>
> **Answer c:**
>
> ionic
>
> **Answer d:**

> not ionic

The Crystaline Structure of Ionic Compounds

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure 6.4.1). 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 6.4.1 shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.
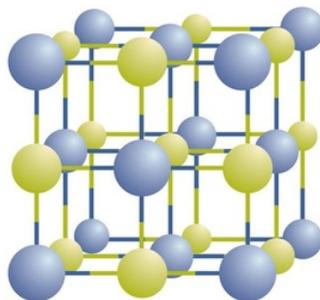
Figure 6.4.1: A Sodium Chloride Crystal. A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

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:

$$Mg^{2+}Cl^{-} \tag{6.4.1}$$
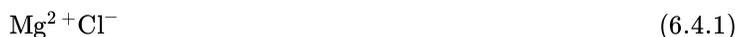
Combining one ion of each does not completely balance the positive and negative charges. The easiest way to balance these charges is to assume the presence of *two* chloride ions for each magnesium ion:

$$Mg^{2+}Cl^{-}Cl^{-} \tag{6.4.2}$$

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.

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.
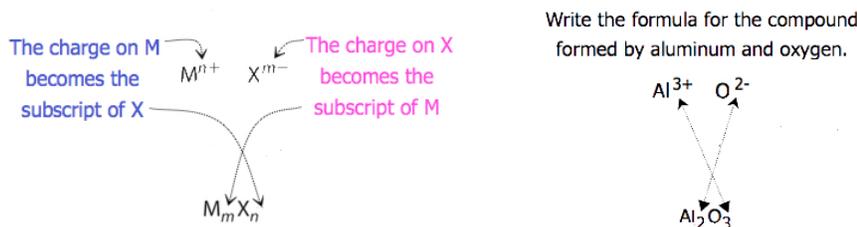
Figure 6.4.2: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges.

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 6.4.1**

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. 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 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 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 Fe_2O_3. Alternatively, use the crossing charges method shown in Figure 3.3.2.

? **Exercise 6.4.1**

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

Formula that Contain Polyatomic Ions

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:

Write the formula for the compound formed by

a. barium and nitrate   b. ammonium and phosphate

Ba^{2+} NO_3^- NH_4^+ PO_4^{3-}

$Ba(NO_3)_2$ $(NH_4)_3PO_4$
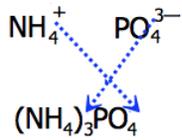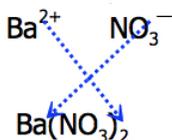
✔ **Example** 6.4.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. a potassium ion and a sulfate ion
b. a calcium ion and a nitrate ion
c. a magnesium ion and a hydroxide 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.
c. Magnesium ions have a +2 charge and hydroxide ions have a -1 charge. Two hydroxide ions are therefore needed to balance one magnesium ion. The formula for hydroxide must be in parenthesis because two of them appear in the formula. The chemical formula is $Mg(OH)_2$

❓ **Exercise** 6.4.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. a magnesium ion and a carbonate ion
b. an aluminum ion and an acetate ion
c. a zinc ion and a cyanide ion

Answer a:
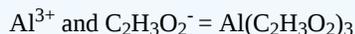
Mg^{2+} and CO_3^{2-} = $MgCO_3$

Answer b:

Al^{3+} and $C_2H_3O_2^-$ = $Al(C_2H_3O_2)_3$

Answer c:

Zn^{2+} and CN^- = $Zn(CN)_2$

6.5: Naming Ionic Compounds

> 🧠 **Learning Objectives**
>
> - Given the formula of an ionic compound (which may contain a fixed-charge or variable-charge metal and may contain a polyatomic ion) write the name.

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 6.5.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

Name of cation (metal) + Base name of anion (nonmetal) and *-ide*

$BaCl_2$

Ba^{2+} Cl^-
                   Cl^-

barium ion   chloride ion

Remove the word "ion"

barium + chloride
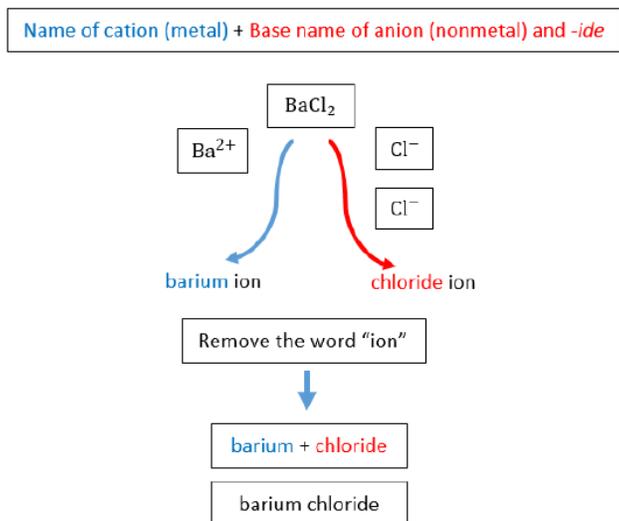
barium chloride

Figure 6.5.1: Naming $BaCl_2$

Subscripts in the formula do not affect the name.

> ✔ **Example 6.5.3: Naming Ionic Compounds**
>
> Name each ionic compound.
>
> a. $CaCl_2$
> b. AlF_3
> c. KCl
>
> **Solution**
>
> a. Using the names of the ions, this ionic compound is named calcium chloride.
> b. The name of this ionic compound is aluminum fluoride.
> c. The name of this ionic compound is potassium chloride

> ❓ **Exercise 6.5.3**
>
> Name each ionic compound.
>
> a. AgI
> b. MgO
> c. Ca_3P_2
>
> **Answer a:**
>   silver iodide

Answer b:

　　magnesium oxide

Answer c:

　　calcium phosphide

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

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. The names are iron(II) chloride and iron(III) chloride (Figure $6.5.2$).

Table 6.5.3: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.

Name of cation (metal) + (Roman Numeral in parenthesis) + Base name of anion (nonmetal) and -ide
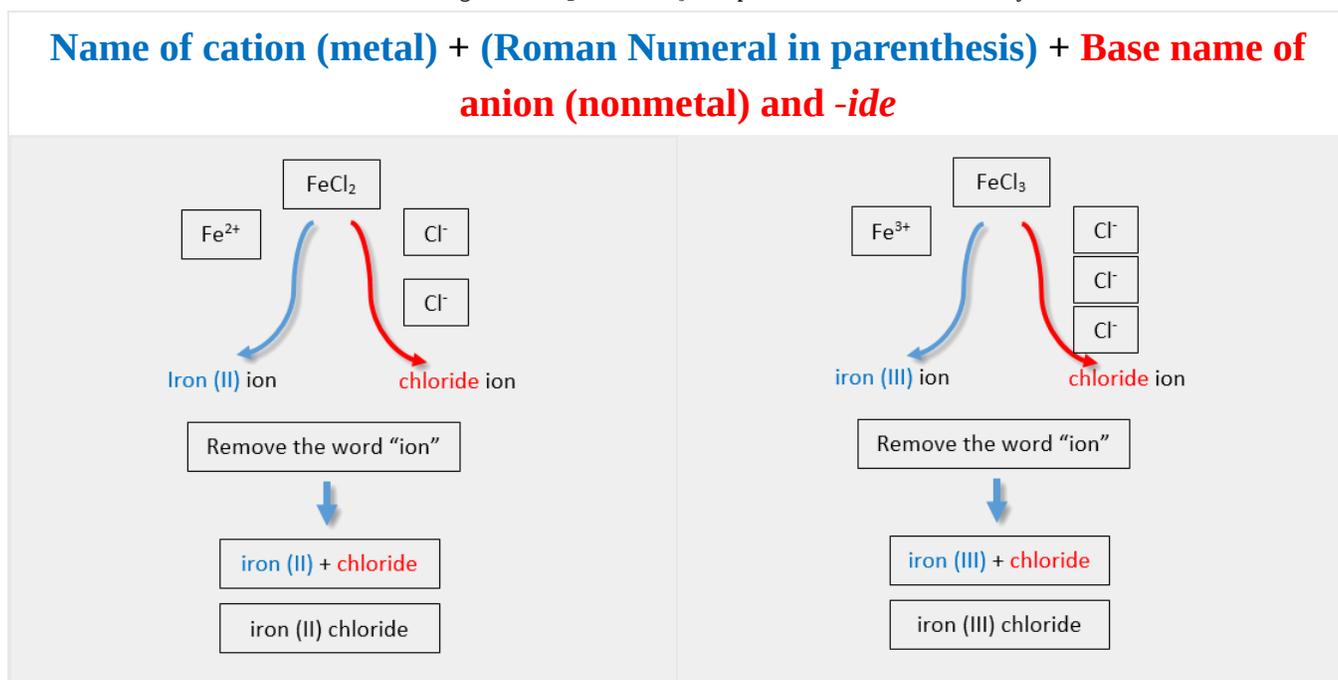
Figure $6.5.2$: Naming Ionic Compounds with Variable Charge Metals

---

✔ Example 6.5.4:

Name each ionic compound.

　a. Co_2O_3
　b. $FeCl_2$

Solution

Explanation	Answer

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloride always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−. • This means that the one iron ion must have a 2+ charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

? Exercise 6.5.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium sulfate compound in Figure 6.5.6.
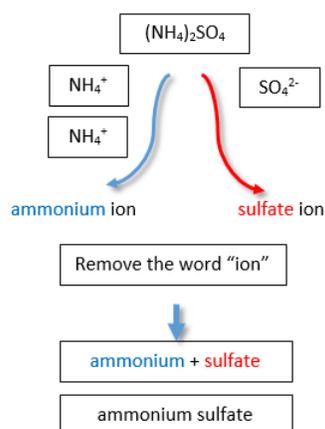
Figure 6.5.3: Naming Ionic Compounds with Polyatomic Ions

✔ Example 6.5.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound's name is iron(II) phosphate.	iron(II) phosphate

---

? Exercise 6.5.5A

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:
   ammonium phosphate

Answer b:
   cobalt(III) nitrite

Figure 6.5.1 is a synopsis of how to name simple ionic compounds.
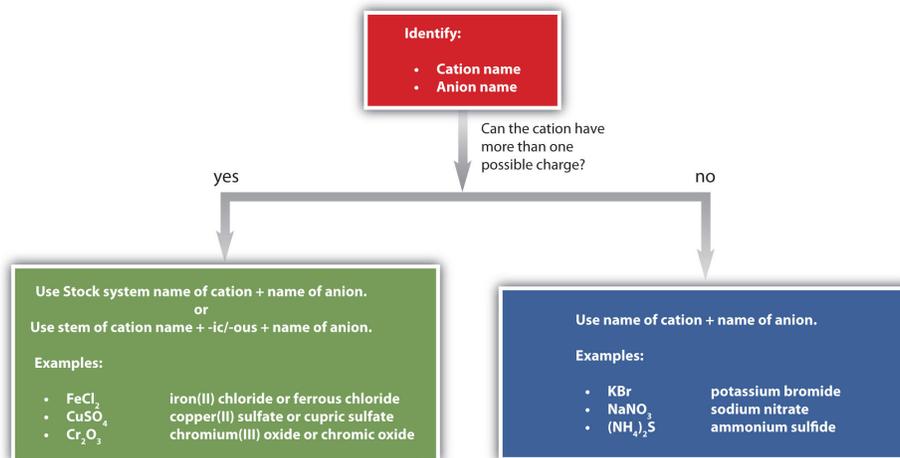
Figure 6.5.3: A Guide to Naming Simple Ionic Compounds.

❓ Exercise 6.5.5B

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:
    zinc bromide

Answer b:
    aluminum oxide

Answer c:
    ammonium phosphate

Answer d:
    gold(III) fluoride

Answer e:
    silver fluoride

Contributions & Attributions

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7: Molecules, Covalent Bonding, and the Nomenclature of Binary Covalent Compounds

7.1: Covalent Bonds and Molecules

> ### ◑ Learning Objectives
>
> - Describe covalent compounds as existing as discrete molecules.
> - Explain that the formula of a covalent compound specifies the number of atoms in the molecule.

You have already seen examples of substances that contain covalent bonds. One substance mentioned previously was water (H_2O). You can tell from its formula that it is not an ionic compound; it is not composed of a metal and a nonmetal. Consequently, its properties are different from those of ionic compounds.

Electron Sharing

Previously, we discussed ionic bonding in which atoms gain or lose electrons and the resulting positively charged cations and negatively charged anions become attracted to one another. **Covalent bonds**, on the other hand, form from the *sharing* of electrons between atoms. This concept can be illustrated by using two hydrogen atoms. We can represent the two individual hydrogen atoms as follows:
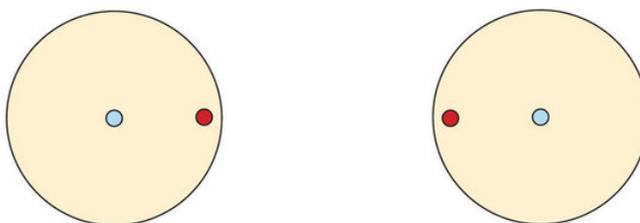
In contrast, when two hydrogen atoms get close enough together to share their electrons, they can be represented as follows:
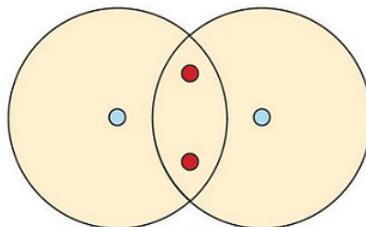
By sharing their electrons, both hydrogen atoms now have two electrons. This arrangement is more stable than when the two atoms each have a single electron. The sharing of electrons between atoms is called a **covalent bond**, and the two electrons that join atoms in a covalent bond are called a bonding pair of electrons. A discrete group of atoms connected by covalent bonds is called a **molecule**—the smallest part of a covalent compound that retains the chemical identity of that compound.

Chemists frequently use Lewis diagrams to represent covalent bonding in molecular substances. For example, the Lewis diagrams of two separate hydrogen atoms are as follows:

$$H \cdot \qquad \cdot H$$

The Lewis diagram of two hydrogen atoms sharing electrons looks like this:

$$H : H$$

This depiction of molecules is simplified further by using a dash to represent a covalent bond. The hydrogen molecule is then represented as follows:

$$H - H$$
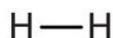
Remember that the dash, also referred to as a single bond, represents a *pair* of electrons.

The bond in a hydrogen molecule, measured as the distance between the two nuclei, is about 7.4×10^{-11} m, or 74 picometers (pm; 1 pm = 1×10^{-12} m). This particular bond length represents a balance between several forces: the attractions between oppositely

charged electrons and nuclei, the repulsion between two negatively charged electrons, and the repulsion between two positively charged nuclei. If the nuclei were closer together, they would repel each other more strongly; if the nuclei were farther apart, there would be less attraction between the positive and negative particles.

The nature of covalent bonding and of molecules is the overriding subject of the first term of General Chemistry (CH221). For now, it is sufficient to know the following points:

- Covalent bonding occurs primarily between nonmetal atoms. A metalloid atom paired with a nonmetal atom can also form a covalent bond.
- Covalent bonds result from two atoms sharing one or more pairs of electrons.
- When a covalent bond forms, the resulting molecule is more stable than the atoms it came from.
- The two atoms participating in a covalent bond become geometrically fixed in space relative to one another.
- Two or more atoms covalently bonded together from a molecule.
- Lewis structures are used to represent molecules. How to draw and interpret Lewis structures will be covered in CH221.

7.2: Contrasting Ionic Compounds and Covalent Compounds

> **⬤ Learning Objectives**
>
> - Describe ionic compounds as an extended three dimensional array, lattice structures, versus discrete molecules in covalent compounds.
> - Explain that the formula of an ionic compound specifies the ratio of cations to anions, while the molecular formula specifies that number of atoms in the molecule.

We have seen that ionic and molecular compounds differ greatly from each other on the particle level. These structural differences lead to very different physical properties on the macroscopic level.

Physical Properties of Molecular Compounds

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. **Covalent compounds**, sometimes called molecular compounds, display a wide range of physical properties due to the many different sizes, shapes, and compositions of molecules. The melting and boiling points of covalent compounds are generally quite low compared to those of **ionic compounds**. This is because melting of ionic compounds involves breaking ionic bonds whereas the melting of covalent compounds involves disrupting the weak forces between molecules. When covalent compounds melt, the covalent bonds in the molecules are not broken.

Since covalent compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. Ionic compounds do not conduct electricity in the solid state because of their rigid structure, but conduct well when either molten or dissolved into a solution. The water solubility of covalent compounds is variable and depends on the structure of the molecule. Many, but not all, ionic compounds are quite soluble in water. The table below summarizes some of the differences between ionic and covalent compounds.

Also note that the chemical formula of a covalent compound represents something slightly different than the chemical formula of an ionic compound. The chemical formula of a covalent compound represents the identity and actual number of atoms that compose a molecule. The chemical formula of an ionic compound, on the other hand, represents the simplist whole number ratio of anions to cations in the crystal.

Property	Ionic Compounds	Covalent Compounds
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - attraction between anions and cations	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
What the formula represents	The ratio of cations to anions	The type and number of atoms in the molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor
State when dissolved in water	Separates into ions	Remains molecules

Table 5.8.1: Comparison of Ionic and Molecular Compounds

7.3: The Dissolving Process- Ionic Compounds Versus Covalent Compounds

> ◐ **Learning Objectives**
>
> - Differentiate between aqueous solutions of ionic and covalent compounds.

The Process of Dissolution

Water dissolves many ionic compounds and some covalent compounds. The mechanism by which ionic compounds and covalent compounds dissolve, however, is different.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative ions. Water is attracted to the sodium chloride crystal because water has both a slightly positive end (the hydrogens) and a sightly negative end (the oxygen). The positively charged sodium ions in the crystal attract the oxygen end of the water molecules. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules. The action of the water molecules takes the crystal lattice apart (see image below).
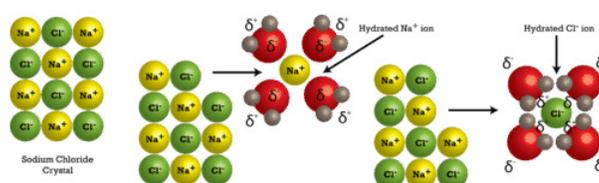
Figure $7.3.1$: The dissolving of sodium chloride in water.

After coming apart from the crystal, the individual ions are then surrounded by water molecules. Note that the individual Na^+ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. Ions being surrounded by water molecules helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and reforming a crystal.

Table sugar is sucrose $(C_{12}H_{22}O_{11})$, and is an example of a covalent compound. Solid sugar consists of individual sugar molecules held together by relatively weak forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the relatively weak forces between molecules, but it does not break the covalent bonds between the carbon, hydrogen, and oxygen atoms within the molecule. The attractive forces between water molecules and solute molecules are also different than the attractive forces between water molecules and hydrated ions. This topic will be discussed in more depth in CH222.

7.4: Binary Covalent Compounds- Formulas and Names

> **◑ Learning Objectives**
>
> - Given the formula of a binary covalent inorganic compound, write the systematic name.
> - Given the systematic name for a binary covalent inorganic compound, write the 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-}.

ammonium ion   carbonate ion
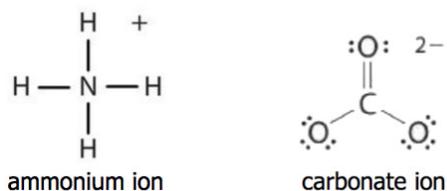
The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. Nonmetal atoms in polyatomic ions are joined by covalent bonds, but the ion as a whole participates in ionic bonding. For example, ammonium chloride (NH_4Cl) has ionic bonding between a polyatomic 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).

> **✔ Example 7.4.1**
>
> 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 7.4.1

Is each compound are formed from ionic bonds, covalent bonds, or both?

a. $Ba(OH)_2$
b. F_2
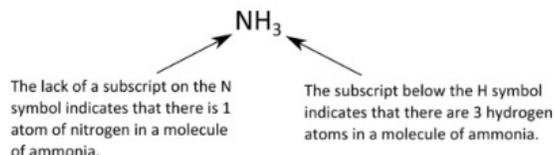c. PCl_3

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 and acids are the prominent exceptions). 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.

The lack of a subscript on the N symbol indicates that there is 1 atom of nitrogen in a molecule of ammonia.

The subscript below the H symbol indicates that there are 3 hydrogen atoms in a molecule of ammonia.

Unlike ionic compounds, the subscripts for covalent compounds are never reduced. For example, a molecule of hydrogen peroxide contains two oxygen atoms and two hydrogen atoms. This means the formula is H_2O_2. it would be incorrect to reduce it to HO since this would imply that the molecule contains only one hydrogen atom and one oxygen atom.

Naming Covalent Compounds

Naming *binary* (two-element) covalent compounds is different than naming 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*. (Note that the -ide suffix here does not indicate the presence of a monatomic anion.) A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 7.4.1 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 7.4.1: 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-

*This prefix is not used for the first element's name.

Number of Atoms in Compound	Prefix on the Name of the Element
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 7.4.2

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 7.4.2

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

✔ Example 7.4.3

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 7.4.3

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 some of these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

7.4: Binary Covalent Compounds- Formulas and Names is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by LibreTexts.

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7.5: Drawing Lewis Structures

> ⚓ **Learning Objectives**
>
> - Given the molecular formula of a covalent compound or polyatomic ion students will be able to draw a Lewis structure.

Drawing Lewis Structures

For very simple molecules and polyatomic ions, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms. See these examples:

$$H\cdot \ + \ :\ddot{B}r\cdot \ \longrightarrow \ H:\ddot{B}r:$$

$$2H\cdot \ + \ :\ddot{S}\cdot \ \longrightarrow \ H:\ddot{S}: \\ H$$

$$\cdot\ddot{N}\cdot \ + \ \cdot\ddot{N}\cdot \ \longrightarrow \ :N:::N:$$
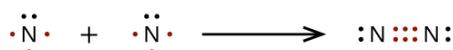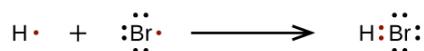
For more complicated molecules and polyatomic ions, it is helpful to follow the step-by-step procedure outlined here:

1. Determine the total number of valence (outer shell) electrons among all the atoms. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.
2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom. (Generally, the least electronegative element should be placed in the center.) Connect each atom to the central atom with a single bond (one electron pair).
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen), completing an octet around each atom.
4. Place all remaining electrons on the central atom.
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

Let us determine the Lewis structures of OF_2 and HCN as examples in following this procedure:

1. Determine the total number of valence (outer shell) electrons in the molecule or ion. For a molecule, we add the number of valence electrons (use the main group number) on each atom in the molecule. This is the total number of electrons that must be used in the Lewis structure.

$$O + 2\,(F) = OF_2 \qquad H + C + N = HCN$$

$$6e^- + (2 \text{ x } 7e^-) = 20e^- \qquad 1e^- + 4e^- + 5e^- = 10e^-$$

2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom and connecting each atom to the central atom with a single (one electron pair) bond. Note that H and F can only form one bond, and are always on the periphery rather than the central atom.

$$F—O—F \qquad H—C—N$$

3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons.

- In OF_2, six electrons are placed on each F.
- In HCN, six electrons placed on N

$$:\ddot{F}—O—\ddot{F}: \qquad H—C—\ddot{N}:$$

4. Place all remaining electrons on the central atom.

- In OF_2, 4 electrons are placed on O.
- In HCN: no electrons remain (the total valence of $10e^-$ is reached) so nothing changes.
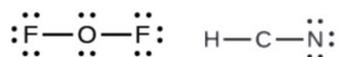
$$\ddot{:}\!F\!-\!\ddot{O}\!-\!\ddot{F}\!: \qquad H\!-\!C\!-\!\ddot{N}\!:$$

5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

- In OF_2, each atom has an octet as drawn, so nothing changes.
- In HCN, form two more C–N bonds

$$H\!-\!C\!-\!\ddot{N}\!:$$
$$\text{gives}$$
$$\ddot{:}\!F\!-\!\ddot{O}\!-\!\ddot{F}\!: \qquad H\!-\!C\!\equiv\!N\!:$$
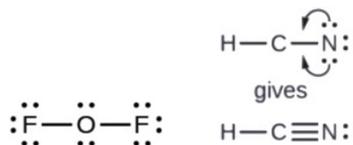
Finally, check to see if the total number of valence electrons are present in the Lewis structure. And then, inspect if the H atom has 2 electrons surrounding it and if each of the main group atoms is surrounded by 8 electrons.

Multiple Bonds

In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares only two electrons. Review HCN in Step 5 above. Another example is carbon dioxide (CO_2). CO_2 has a total valence of $4e^- + (2 \times 6e^-) = 16e^-$. Following steps 1 to 4, we draw the following:

$$\text{only } 4\,e^-$$
$$\downarrow$$
$$\ddot{:}\!O\!:\!)C(\!:\!O\!:$$
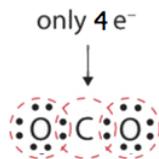
This does not give the carbon atom a complete octet; only four electrons are in its valence shell. This arrangement of shared electrons is far from satisfactory.

$$\ddot{:}\!O\!:\!)C(\!:\!O\!:$$

In this case, more than one pair of electrons must be shared between two atoms for both atoms to have an octet. A second electron pair from each oxygen atom must be shared with the central carbon atom shown by the arrows above. A lone pair from each O must be converted into a bonding pair of electrons.

$$\ddot{:}\!O\!::\!)C(\!::\!O\!:$$
$$\uparrow\uparrow\uparrow$$
$$8e^-$$
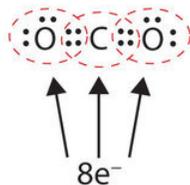
In this arrangement, the carbon atom shares four electrons (two pairs) with the oxygen atom on the left and four electrons with the oxygen atom on the right. There are now eight electrons around each atom. Two pairs of electrons shared between two atoms make a double bond between the atoms, which is represented by a double dash:

$$\ddot{:}\!O\!=\!C\!=\!\ddot{O}\!:$$

Some molecules contain triple bonds (like HCN, shown above). Triple bonds are covalent bonds in which *three* pairs of electrons are shared by two atoms. Another compound that has a triple bond is acetylene (C_2H_2), whose Lewis diagram is as follows:

$$H:C:::C:H \quad \text{or} \quad H:C \equiv C:H$$

✔ **Example 7.5.1**

Draw the Lewis diagram for each molecule.

1. N_2
2. CH_2O (The carbon atom is the central atom.) One application of CH_2O, also called formaldehyde, is the preservation of biological specimens. Aqueous solutions of CH_2O are called formalin and have a sharp, characteristic (pungent) odor.

Solution

1. The total number of electrons is 2 x 5 = 10 electrons. The bond between the two nitrogen atoms is a triple bond. The Lewis diagram for N_2 is as follows:

$$:N:::N: \qquad\qquad \text{or} \qquad\qquad :N \equiv N:$$

2. The total number of electrons is 4 x 2(1) + 6 = 12 electrons. In CH_2O, the central atom is surrounded by two different types of atoms. The Lewis diagram that fills each atom's valence electron shell is as follows:

? **Exercise 7.5.1**

Draw the Lewis diagram for each molecule.

a. O_2
b. C_2H_4

Answer a:

Answer b:

✔ **Example 7.5.2**

Which is the correct Lewis structure for N_2H_2?

A. $H - \ddot{N} = \ddot{N} - H$

B. $H - N \equiv N - H$

C. $H - \ddot{N} \equiv \ddot{N} - H$

Solution

Lewis structure A is the correct answer. It has a total of $(2 \times 5e^-) + (2 \times 1e^-) = 12e^-$. Each of the N atoms satisfy the octet requirement and the H atoms follow the duet rule.

Structure B is electron deficient. It has only $10e^-$ instead of 12.

Structure C has 14 (2 extra) electrons. The N atoms do not satisfy the octet.

Exercise 7.5.2

Which is the correct Lewis structure for NOCl?

A. $:\ddot{\text{O}}-\ddot{\text{N}}-\ddot{\text{C}}\text{l}:$

B. $:\ddot{\text{O}}-\ddot{\text{N}}=\ddot{\text{C}}\text{l}:$

C. $\ddot{\text{O}}=\ddot{\text{N}}-\ddot{\text{C}}\text{l}:$
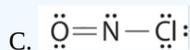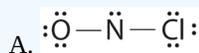
Answer

Structure A violates the octet rule; N is surrounded by only $6e^-$.

Structure B violates the octet rule; Cl has $10e^-$ around it. Furthermore, there are a total of $20e^-$ instead of $18e^-$.

Structure C is the correct structure. It has a total of $6e^- + 5e^- + 7e^- = 18e^-$. Each atom is surrounded by 8 electrons (octet rule).

Exercises

1. What is one clue that a molecule has a multiple bond?

2. Draw the Lewis diagram for each of the following.

    a. H_2O

    b. NH_3

    c. C_2H_6

    d. CCl_4

3. Each molecule contains double bonds. Draw the Lewis diagram for each. The first element is the central atom.

    a. CS_2
    b. C_2F_4
    c. $COCl_2$

4. Each molecule contains multiple bonds. Draw the Lewis diagram for each. Assume that the first element is the central atom, unless otherwise noted.

    a. N_2
    b. CO
    c. HCN (The carbon atom is the central atom.)
    d. POCl (The phosphorus atom is the central atom.)

5. Explain why hydrogen atoms do not form double bonds.

Answers

1. If single bonds between all atoms do not give all atoms (except hydrogen) an octet, multiple covalent bonds may be present.
2. a.

$\text{H}-\ddot{\underset{..}{\text{O}}}-\text{H}$

b.

H—N—H
   |
   H

c.

     H H
     | |
H—C—C—H
     | |
     H H

d.

        :Cl:
          |
:Cl—C—Cl:
          |
        :Cl:

3. a.

:S=C=S:

b.

:F:C=C:F:
  :F: :F:

c.

      :O:
       :C:
  :Cl: :Cl:

4. a. N≡N

b. :C≡O:

c. H:C≡N

d.

     P=O:
      /
  :Cl:
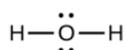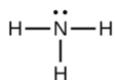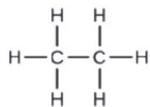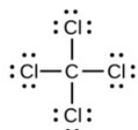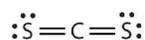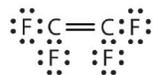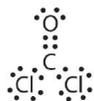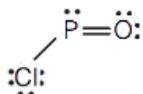
5. Hydrogen can accept only one more electron; multiple bonds require more than one electron pair to be shared.

7.6: Resonance

> **◑ Learning Objectives**
>
> - Given the molecular formula of a covalent compound or polyatomic ion students will be able to draw a Lewis structure.

Resonance describes the delocalization of electrons within molecules. It involves constructing multiple Lewis structures that, when combined, represent the full electronic structure of the molecule. **Resonance structures** are used when a single Lewis structure cannot fully describe the bonding; the combination of possible resonance structures is defined as a **resonance hybrid**, which represents the overall delocalization of electrons within the molecule. In general, molecules with multiple resonance structures will be more stable than one with fewer.

Introduction

Resonance is a way of describing delocalized electrons within certain molecules or polyatomic ions where the bonding cannot be expressed by a single Lewis formula. A molecule or ion with such delocalized electrons is represented by several resonance structures. The nuclear skeleton of the Lewis Structure of these resonance structures remains the same, only the electron locations differ. Such is the case for ozone (O_3), an allotrope of oxygen.

1. We know that ozone has a three atoms, so one O atom is central and the other two are bonded to it:

$$O$$
$$O \quad O$$

2. Each O atom has 6 valence electrons, for a total of 18 valence electrons.

3. Assigning one bonding pair of electrons to each oxygen–oxygen bond gives

with 14 electrons left over.

4. If we place three lone pairs of electrons on each terminal oxygen, we obtain
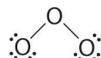
and have 2 electrons left over.

5. At this point, both terminal oxygen atoms have octets of electrons. We therefore place the last 2 electrons on the central atom:
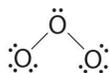
6. The central oxygen has only 6 electrons. We must convert one lone pair on a terminal oxygen atom to a bonding pair of electrons—but which one? Depending on which one we choose, we obtain either

     or
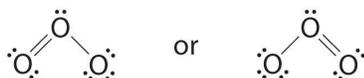
Which is correct? In fact, neither is correct. Both predict one O–O single bond and one O=O double bond. If the bonds were of different types (one single and one double, for example), they would have different lengths. It turns out, however, that both O–O bond distances are identical, 127.2 pm, which is shorter than a typical O–O single bond (148 pm) and longer than the O=O double bond in O_2 (120.7 pm).

Equivalent Lewis dot structures, such as those of ozone, are called **resonance structures**. The position of the *atoms* is the same in the various resonance structures of a compound, but the position of the *electrons* is different. Double-headed arrows link the

different resonance structures of a compound:
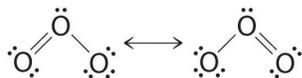
The double-headed arrow indicates that the actual electronic structure is an *average* of those shown, not that the molecule oscillates between the two structures.

> When it is possible to write more than one equivalent resonance structure for a molecule or ion, the actual structure is the average of the resonance structures.

The electrons appear to "shift" between different resonance structures and while not strictly correct as each resonance structure is just a limitation of using the Lewis structure perspective to describe these molecules. A more accurate description of the electron structure of the molecule requires considering multiple resonance structures simultaneously.

📌 Delocalization and Resonance Structures Rules

1. Resonance structures should have the same number of electrons, do not add or subtract any electrons. (check the number of electrons by simply counting them).
2. Each resonance structures follows the rules of writing Lewis Structures.
3. The hybridization of the structure must stay the same.
4. The skeleton of the structure can not be changed (only the electrons move).
5. Resonance structures must also have the same number of lone pairs.

✏️ "Pick the Correct Arrow for the Job"

Most arrows in chemistry cannot be used interchangeably and care must be given to selecting the correct arrow for the job.

- ↔: A double headed arrow on both ends of the arrow between Lewis structures is used to show resonance
- ⇌: Double harpoons are used to designate equilibria
- ⇀: A single harpoon on one end indicates the movement of **one** electron
- →: A double headed arrow on one end is used to indicate the movement of **two** electrons

✔ Example 7.6.1: Carbonate Ion

Identify the resonance structures for the carbonate ion: CO_3^{2-}.

Solution

1. Because carbon is the least electronegative element, we place it in the central position:

2. Carbon has 4 valence electrons, each oxygen has 6 valence electrons, and there are 2 more for the −2 charge. This gives 4 + (3 × 6) + 2 = 24 valence electrons.

3. Six electrons are used to form three bonding pairs between the oxygen atoms and the carbon:

4. We divide the remaining 18 electrons equally among the three oxygen atoms by placing three lone pairs on each and indicating the −2 charge:
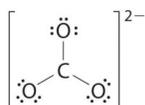
5. No electrons are left for the central atom.

6. At this point, the carbon atom has only 6 valence electrons, so we must take one lone pair from an oxygen and use it to form a carbon–oxygen double bond. In this case, however, there are *three* possible choices:
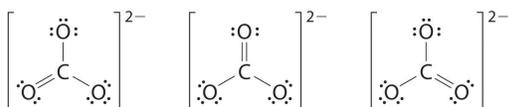
As with ozone, none of these structures describes the bonding exactly. Each predicts one carbon–oxygen double bond and two carbon–oxygen single bonds, but experimentally all C–O bond lengths are identical. We can write resonance structures (in this case, three of them) for the carbonate ion:
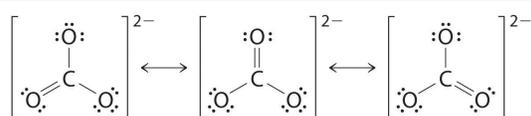
The actual structure is an average of these three resonance structures.

Like ozone, the electronic structure of the carbonate ion cannot be described by a single Lewis electron structure. Unlike O_3, though, the actual structure of CO_3^{2-} is an average of *three* resonance structures.

✔ Example 7.6.2: Nitrate Ion

Draw the possible resonance structures for the Nitrate ion NO_3^-.

Solution

1. Count up the valence electrons: $(1*5) + (3*6) + 1(\text{ion}) = $ **24** electrons

2. Draw the bond connectivities:
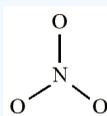
3. Add octet electrons to the atoms bonded to the center atom:
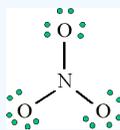
4. Place any leftover electrons (24-24 = **0**) on the center atom:
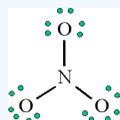
5. Does the central atom have an octet?

- **NO**, it has 6 electrons
- Add a multiple bond (first try a double bond) to see if the central atom can achieve an octet:
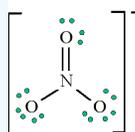
6. Does the central atom have an octet?

- YES
- Are there possible resonance structures? YES
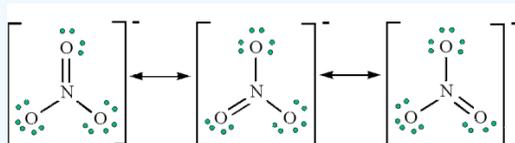
Note: We would expect that the bond lengths in the NO_3^- ion to be somewhat shorter than a single bond.

References

1. Petrucci, Ralph H., et al. *General Chemistry: Principles and Modern Applications*. New Jersey: Pearson Prentice Hall, 2007.
2. Ahmad, Wan-Yaacob and Zakaria, Mat B. "Drawing Lewis Structures from Lewis Symbols: A Direct Electron Pairing Approach." Journal of Chemical Education: Journal 77.3.

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7.7: Molecular Geometry- VSEPR

> **◖◗ Learning Objectives**
>
> - Given the molecular formula of a covalent compound or polyatomic ion students will be able to determine the Electron and Molecular Geometry.

Unlike the ions in ionic compounds, which are arranged in lattices called crystals, molecules of covalent compounds exist as discrete units with a certain three-dimensional shape.

Molecular Shape: VSEPR Theory

Unlike ionic compounds, with their extended crystal lattices, covalent molecules are discrete units with specific three-dimensional shapes. The shape of a molecule is determined by the fact that covalent bonds, which are composed of negatively charged electrons, tend to repel one another. This concept is called the **valence shell electron pair repulsion (VSEPR) theory**. For example, the two covalent bonds in $BeCl_2$ **stay as far from each other** as possible, ending up 180° apart from each other. The result is a *linear* molecule:
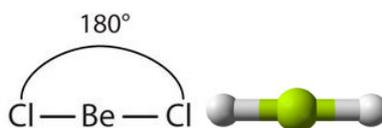
The three covalent bonds in BF_3 repel each other to form 120° angles in a plane, in a shape called *trigonal planar*:
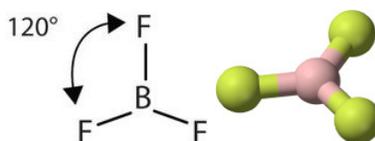
The molecules $BeCl_2$ and BF_3 actually violate the octet rule; however, such exceptions are rare.

> Try sticking three toothpicks into a marshmallow or a gumdrop and see if you can find different positions where your "bonds" are farther apart than the planar 120° orientation.

The four covalent bonds in CCl_4 arrange themselves three dimensionally, pointing toward the corner of a tetrahedron and making bond angles of 109.5°. CCl_4 is said to have a *tetrahedral* shape:
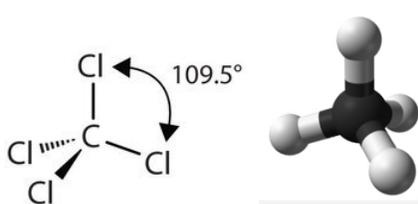
Atoms Around Central Atom	Geometry	Example
2 AB_2	Linear	$BeCl_2$
3 AB_3	Trigonal Planar	BF_3
4 AB_4	Tetrahedral	CCl_4

In determining the shapes of molecules, it is useful to first determine the Lewis diagram for a molecule. The shapes of molecules with multiple bonds are determined by treating the multiple bonds as one bond. Thus, formaldehyde (CH_2O) has a shape similar to that of BF_3. It is *trigonal planar*.
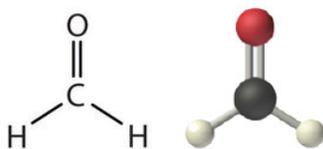
It is often useful to talk about both the **electron geometry** of the molecule and the **molecular geometry**. **Electron geometry** looks at the arrangement of all electron-rich areas (both lone pairs and covalent bonds) around the central atom. **Molecular geometry** looks at only the arrangement of covalent bonds. For molecules like BF_3, CCl_4, and CH_2O in which the central atom has no lone pairs, the electron geometry and the molecular geometry are the same.

Molecules With Lone Pairs Around Central Atom

Molecules with lone electron pairs around the central atom have a molecular geometry based on the position of the atoms, not the electron pairs. For example, NH_3 has **one lone electron pair** and **three bonded electron pairs**. These four electron pairs repel each other and adopt a tetrahedral electron geometry. However, the molecular geometry of the molecule is described in terms of the positions of the atoms, not the lone electron pairs. Thus, NH_3 is said to have a *trigonal pyramidal* molecular geometry, not a tetrahedral one. Whenever there are lone electrons pairs on the central atom, the electron geometry and molecular geometry will be different.
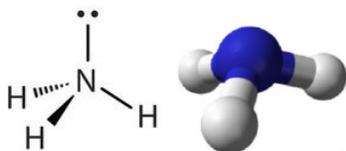
Similarly, H_2O has **two lone pairs** of electrons around the central oxygen atom and **two bonded electron pairs**. Although the four electron geometry is tetrahedral, the molecular geometry of the molecule is described by the positions of the atoms only. The molecular geometry of H_2O is *bent* with an approximate 109.5° angle.
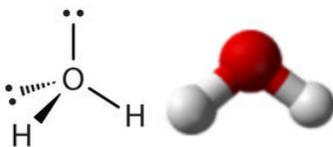
In summary, to determine the molecular geometry:

Step 1: Draw the Lewis structure.

Step 2: Count the number of bonds (a double/triple bond counts as one) and lone pairs around the **central atom**.

Step 3: Use Table 4.5.1 to determine the molecular geometry.

Number of bonds and lone pairs	Number of bonds 0 lone pair	1 lone pair	2 lone pairs
2	180° X—A—X Linear		
3	X 120° A X X Trigonal planar	A X <120° X Bent or angular	Lone Pair
4	X 109° X A X X Tetrahedral	A X X X<109° Trigonal pyramid	A X X <<109° Bent or angular

*Table 4.5.1:* The molecular geometry depends on the number of bonds and lone pairs around the **central atom, A.**

✔ **Example 7.7.1**

What is the electron geometry and molecular geometry of the ammonium ion, NH_4^+? Its Lewis structure is shown below. How is this different from ammonia, NH_3?

$$\left[\begin{array}{c} H \\ | \\ H-N-H \\ | \\ H \end{array} \right]^+$$
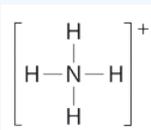
Solution

In ammonium ion, the central atom N has 4 bonds and no lone pairs. It is equivalent to the square shown below from Table 4.5.1. Hence, both the electron geometry and the molecular geometry are *tetrahedral*. Notice that determining the shape of a polyatomic ion is the same as determining the shape of a neutral molecule.

X 109° X A X X
Tetrahedral
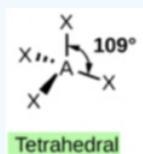
In ammonia (NH_3), shown below, N has 3 bonds and one lone pair.

H—N—H
|
H
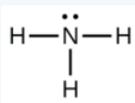
It is equivalent to the square shown below from Table 4.5.1. Hence, the molecular geometry of this molecule is *trigonal pyramidal*. The electron geometry, however, is *tetrahedral*.

A X X X <109°
Trigonal pyramid

Exercise 7.7.1

What is the electron geometry and molecular geometry of nitrosyl chloride, a highly corrosive, reddish-orange gas? Its Lewis structure is shown below.

$$\ddot{O} = \ddot{N} - \ddot{Cl} :$$
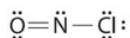
Answer

Focus on the central atom, N. It has a double bond to O, count this as one bond. It also has a single bond to Cl. Thus, N has 2 bonds and one lone pair. These 3 electron pairs will spread out 120 degrees from each other. This means the electron geometry will be *trigonal planar*. But, since the molcular geometry is defined by the arrangement of the atoms only, the shape is bent. If you consult Table 4.5.1, this molecule is equivalent to the below. Hence, two bonds and one lone pair has a *bent* shape.
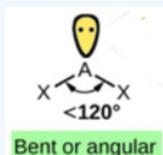
Bent or angular

Exercises

1. What is the electron geometry of each molecule? The molecular geometry?

    a. H_2S
    b. $COCl_2$
    c. SO_2

2. What is the electron geometry of each molecule? The molecular geometry?

    a. NBr_3
    b. SF_2
    c. SiH_4

3. Predict the molecular geometry of nitrous oxide (N_2O), which is used as an anesthetic. A nitrogen atom is in the center of this three-atom molecule.

4. Predict the molecular geometry of acetylene (C_2H_2), which has the two carbon atoms in the middle of the molecule with a triple bond. What generalization can you make about the shapes of molecules that have more than one central atom?

Answers

1. a. electron geometry: tetrahedral, molecular geometry: bent
    b. electron geometry: trigonal planar, molecular geometry: trigonal planar
    c. electron geometry: trigonal planar, molecular geometry: bent

2. a. electron geometry: tetrahedral, molecular geometry: trigonal pyramidal
    b. electron geometry: tetrahedral, molecular geometry: bent
    c. electron geometry: tetrahedral, molecular geometry: tetrahedral

3. linear

4. linear; in a molecule with more than one central atom, the geometry around each central atom needs to be examined.

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7.8: Polarity of Bonds

- Given a table of electronegativity values, students will be able to determine if a covalent bond is polar or nonpolar.

Electronegativity and Bond Polarity

Although we defined covalent bonding as electron sharing, the electrons in a covalent bond are not always shared equally by the two bonded atoms. Unless the bond connects two atoms of the same element, as in H_2, there will always be one atom that attracts the electrons in the bond more strongly than the other atom does, as in HCl, shown in Figure 7.8.1. A covalent bond that has an equal sharing of electrons (Figure 7.8.1a) is called a **nonpolar covalent bond**. A covalent bond that has an unequal sharing of electrons, as in Figure 7.8.1b, is called a **polar covalent bond**.
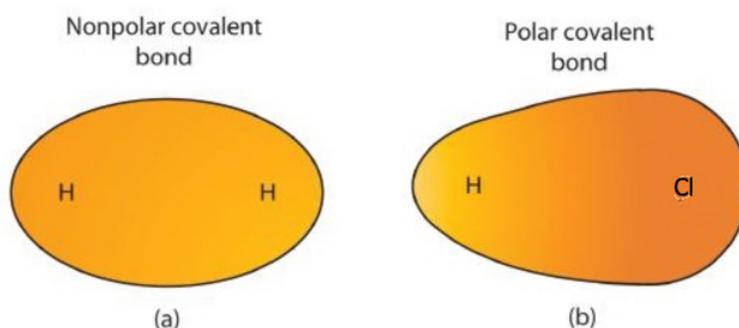
Figure 7.8.1 Polar versus Nonpolar Covalent Bonds. (a) The electrons in the covalent bond are equally shared by both hydrogen atoms. This is a nonpolar covalent bond. (b) The chlorine atom attracts the electrons in the bond more than the hydrogen atom does, leading to an imbalance in the electron distribution. This is a polar covalent bond.

The distribution of electron density in a polar bond is uneven. It is greater around the atom that attracts the electrons more than the other. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom. Note that the shaded area around Cl in Figure 7.8.1b is much larger than it is around H.

This imbalance in electron density results in a buildup of partial negative charge (designated as δ−) on one side of the bond (Cl) and a partial positive charge (designated δ+) on the other side of the bond (H). This is seen in Figure 7.8.2a. The separation of charge in a polar covalent bond results in an electric dipole (two poles), represented by the arrow in Figure 7.8.2b. The direction of the arrow is pointed toward the δ− end while the + tail of the arrow indicates the δ+ end of the bond.
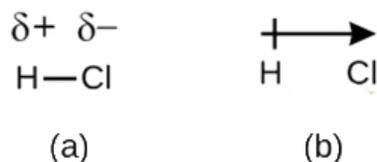
Figure 7.8.2: (a) Unequal sharing of the bonding pair of electrons between H and Cl leads to partial positive charge on the H atom and partial negative charge on the Cl. Symbols δ+ and δ– indicate the polarity of the H–Cl bond. (b) The dipole is represented by an arrow with a cross at the tail. The cross is near the δ+ end and the arrowhead coincides with the δ–.

Any covalent bond between atoms of different elements is a polar bond, but the degree of polarity varies widely. Some bonds between different elements are only minimally polar, while others are strongly polar. Ionic bonds can be considered the ultimate in polarity, with electrons being transferred rather than shared. To judge the relative polarity of a covalent bond, chemists use **electronegativity**, which is a relative measure of how strongly an atom attracts electrons when it forms a covalent bond. There are various numerical scales for rating electronegativity. Figure 7.8.3 shows one of the most popular—the Pauling scale.
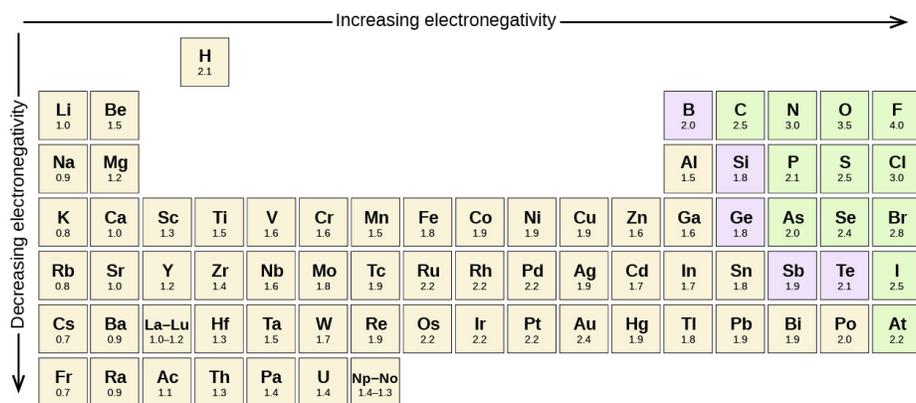
Figure 7.8.3 The electronegativity values derived by Pauling follow predictable periodic trends with the higher electronegativities toward the upper right of the periodic table. Fluorine has the highest value (4.0).

The polarity of a covalent bond can be judged by determining the difference in the electronegativities of the two atoms making the bond. The greater the difference in electronegativities, the greater the imbalance of electron sharing in the bond. Although there are no hard and fast rules, the general rule is if the difference in electronegativities is **less than about 0.4,** the bond is considered **nonpolar;** if the difference is **greater than 0.4,** the bond is considered **polar.** If the difference in electronegativities is large enough (generally **greater than about 1.8),** the resulting compound is considered **ionic** rather than covalent. An electronegativity difference of **zero,** of course, indicates a **nonpolar covalent bond.**

Bond Type	Electronegativity Difference
pure covalent	< 0.4
polar covalent	between 0.4 and 1.8
ionic	> 1.8
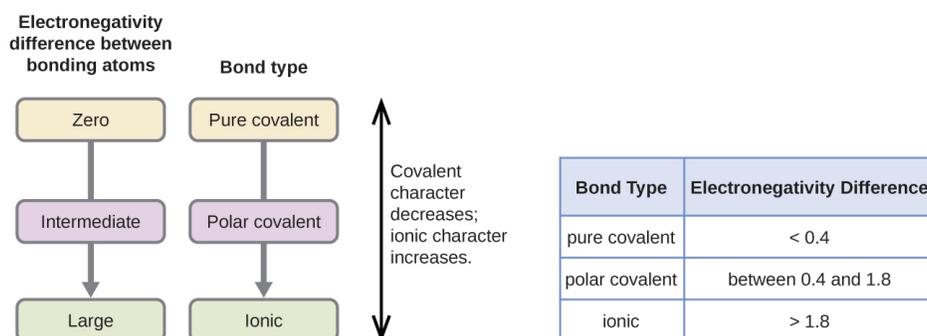
Figure 7.8.4: As the electronegativity difference increases between two atoms, the bond becomes more ionic.

✔ **Example 7.8.1**

Describe the electronegativity difference between each pair of atoms and the resulting polarity (or bond type).

a. C and H
b. H and H
c. Na and Cl
d. O and H

Solution

a. Carbon has an electronegativity of 2.5, while the value for hydrogen is 2.1. The difference is 0.4, which is rather small. The C–H bond is therefore considered nonpolar.
b. Both hydrogen atoms have the same electronegativity value—2.1. The difference is zero, so the bond is nonpolar.
c. Sodium's electronegativity is 0.9, while chlorine's is 3.0. The difference is 2.1, which is rather high, and so sodium and chlorine form an ionic compound.
d. With 2.1 for hydrogen and 3.5 for oxygen, the electronegativity difference is 1.4. We would expect a very polar bond. The sharing of electrons between O and H is unequal with the electrons more strongly drawn towards O.

? Exercise 7.8.1

Describe the electronegativity (EN) difference between each pair of atoms and the resulting polarity (or bond type).

a. C and O
b. K and Br
c. N and N
d. Cs and F

Answer a:

The EN difference is 1.0 , hence polar. The sharing of electrons between C and O is unequal with the electrons more strongly drawn towards O.

Answer b:

The EN difference is greater than 1.8, hence ionic.

Answer c:

Identical atoms have zero EN difference, hence nonpolar.

Answer d:

The EN difference is greater than 1.8, hence ionic.

Exercises

1. Using Figure 7.8.3, determine which atom in each pair has the higher electronegativity.

a. H or C
b. O or Br
c. Na or Rb
d. I or Cl

2. Using Figure 7.8.3, determine which atom in each pair has the lower electronegativity.

a. Mg or O
b. S or F
c. Al or Ga
d. O or I

3. Will the electrons be shared equally or unequally across each covalent bond? If unequally, to which atom are the electrons more strongly drawn?

a. a C–O bond
b. an F–F bond
c. an S–N bond
d. an I–Cl bond

10. Will the electrons be shared equally or unequally across each covalent bond? If unequally, to which atom are the electrons more strongly drawn?

a. a C–C bond
b. a S–Cl bond
c. an O–H bond
d. an H–H bond

11. Arrange the following bonds from least polar to most polar: H-F, H-N, H-O, H-C

12. Arrange the following bonds from least polar to most polar: C-F, C-N, C-O, C-C

Answers

1. Using Figure $7.8.3$, determine which atom in each pair has the higher electronegativity.

 a. C
 b. O
 c. Na
 d. Cl

2. Using Figure $7.8.3$, determine which atom in each pair has the lower electronegativity.

 a. Mg
 b. S
 c. Al
 d. I

3. Will the electrons be shared equally or unequally across each covalent bond? If unequally, to which atom are the electrons more strongly drawn?

 a. unequally toward the O
 b. equally
 c. unequally toward the N
 d. unequally toward the Cl

4. Will the electrons be shared equally or unequally across each covalent bond? If unequally, to which atom are the electrons more strongly drawn?

 a. equally
 b. unequally toward the Cl
 c. unequally toward the O
 d. equally

5. The electronegativity difference increases from 0.4; 0.9; 1.4; 1.9. Hence, the least to most polar: H-C, H-N, H-O, H-F

6. The electronegativity difference increases from 0; 0.5; 1.0; 1.5. Hence, the least to most polar: C-C, C-N, C-O, C-F

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7.9: Polarity of Molecules

> **◀▶ Learning Objectives**
>
> - Given a Lewis structure, students will be able to determine if the molecule is polar or nonpolar.

Molecular Polarity of Diatomic Molecules

If there is only one bond in the molecule, the bond polarity determines the molecular polarity. Any diatomic molecule in which the two atoms are the same element must be a **nonpolar** molecule. A diatomic molecule that consists of a polar covalent bond, such as HF, is a **polar molecule**. A **polar molecule** *is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative*. 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. Hence, a *molecule with two poles is called a* **dipole**. A simplified way to depict polar molecules like HF is pictured below (see figure below).

Figure $7.9.5$: A molecular dipole results from the unequal distribution of electron density throughout a molecule.

When placed between oppositely charged plates, polar molecules orient themselves so that their positive ends are closer to the negative plate and their negative ends are closer to the positive plate (see Figure 4.4.6 below).
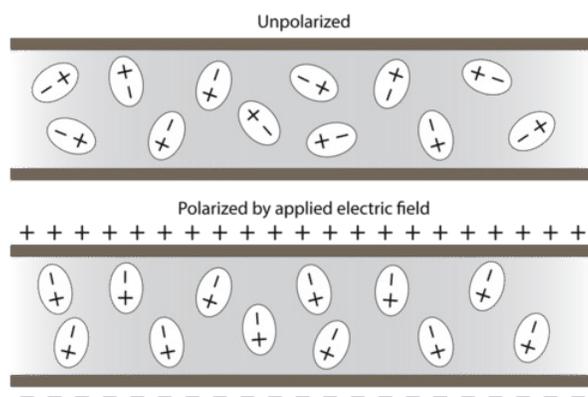
Figure $7.9.6$: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves to maximize the attraction between opposite charges (bottom).

Experimental techniques involving electric fields can be used to determine if a certain substance is composed of polar molecules and to measure the degree of polarity.

Molecular Polarity of Larger Molecules

In general, a molecule is nonpolar if all its bonds are nonpolar. Examples are I_2, O_2, H_2, CH_4, C_2H_6 and C_3H_8.

In general, a molecule is polar if it contains polar bonds EXCEPT when the bond polarities cancel each other. The shape of the CO_2 molecule (linear) orients the two C=O polar bonds **directly opposite** each other, thus **cancelling** each other's effect. **Carbon dioxide** (CO_2) is a *nonpolar* molecule.

On the other hand, water (as discussed above) is a bent molecule because of the two lone pairs on the central oxygen atom. Because of the bent shape, the dipoles do not cancel each other out and the water molecule is *polar*. In the figure below, the individual H-O polar bonds represented by the two red arrows are not directly opposite each other. These two dipoles don't cancel each other out. In fact, the net dipole (blue arrow) points upward. There is a resultant partial positive charge at one end (between the two H atoms) and a partial negative charge on the other end (where O is located). The uneven distribution of charge or the overall dipole is shown by the blue arrow below (Figure 4.5.1). Hence, **water is *polar*** (has + and - poles) while **carbon dioxide is *nonpolar***.
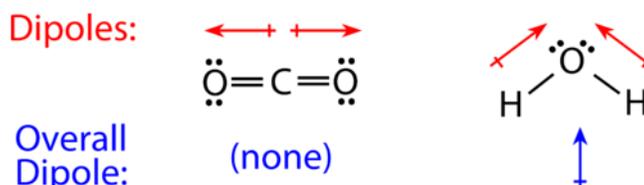
Dipoles:

Overall Dipole: (none)

Figure 4.5.1: *The molecular geometry of a molecule (linear vs. bent) affects its polarity.*

Similarly, in BF_3 (trigonal planar), the effect of a B-F bond is cancelled by the sum of the other two B-F bonds (see video). Hence, a trigonal planar molecule (BF_3) is nonpolar because the bond polarities cancel each other, but a trigonal pyramidal molecule (NH_3) is polar.
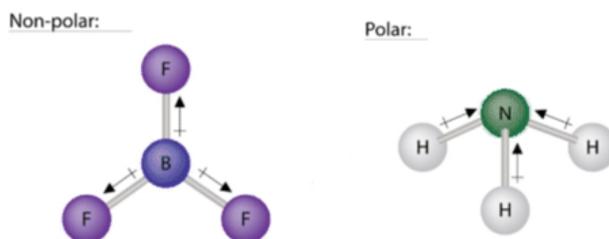
Non-polar: Polar:

Figure 4.5.2: *The molecular geometry of a molecule (trigonal planar vs. trigonal pyramid) affects its polarity.*

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as CCl_4 is ***nonpolar***. However, if the peripheral atoms are not of the same electronegativity, the bond polarities don't cancel and the molecule becomes ***polar***, as in CH_3Cl.
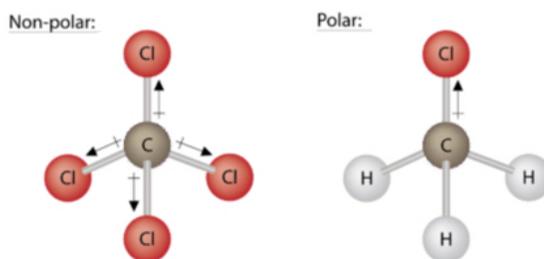
Non-polar: Polar:

Figure 4.5.3: *The same molecular geometry but peripheral bonds are of different electronegativity. CCl_4 is nonpolar but CH_3Cl is polar.*

HCl BCl_3 CH_2O NH_3 $CHCl_3$ CCl_4
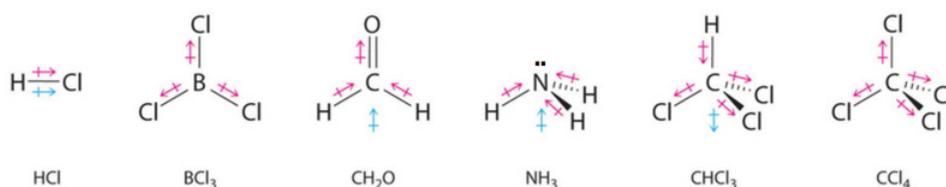
Figure 7.9.4: Molecules with Polar Bonds. Individual bond dipole moments are indicated in red. Due to their different three-dimensional geometry, some molecules with polar bonds have a net dipole moment (HCl, CH_2O, NH_3, and $CHCl_3$), indicated in blue, whereas others do not because the bond dipoles cancel due to symmetry (BCl_3 and CCl_4).

✔ **Example 7.9.2**

Describe the shape of each molecule. Is it polar or nonpolar?

a. PCl_3
b. CO_2

Solution

a. The Lewis diagram for PCl_3 is as follows:

$$:\ddot{Cl} - \overset{..}{\underset{|}{P}} - \ddot{Cl}:$$
$$:\ddot{Cl}:$$
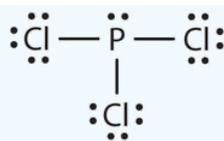
Focus on the central atom, P that has 3 bonds and one lone pair. The four electron pairs arrange themselves tetrahedrally, but the lone electron pair is not considered in describing the molecular shape. Like NH_3, this molecule is *pyramidal*. The 3 P-Cl bonds don't cancel each other. This is *polar*.

b. The Lewis diagram for CO_2 is as follows:

$$:\ddot{O} = C = \ddot{O}:$$
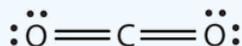
Focus on the central atom, C. The multiple bonds are treated as one group, hence C has 2 bonds and zero lone pair. CO_2 has only two groups of electrons that repel each other. They will direct themselves 180° apart from each other, so CO_2 molecules are *linear*. This is highly symmetrical, with the two opposite dipoles cancelling each other. The CO_2 molecule is *nonpolar*.

? Exercise 7.9.2

Describe the shape of each molecule. Is it polar or nonpolar?

a. CBr_4
b. BCl_3

Answer a:

or
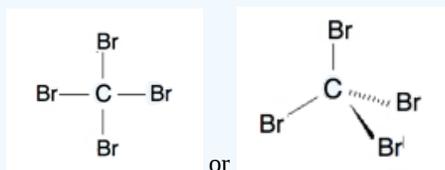
The Lewis structure shows 4 groups attached to the central atom, hence *tetrahedral*. All the 4 groups are identical and the shape is symmetrical. Hence, it is *nonpolar*.

Answer b:
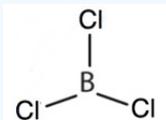
The Lewis diagram shows 3 groups attached to the central atom, hence *trigonal planar*. All the 3 groups are identical and shape is symmetrical, hence, it is *nonpolar*.

Exercises

1. Is each molecule polar or nonpolar? (Reminder: you determined their shapes in the exercises for the last section.)

   a. H_2S
   b. $COCl_2$
   c. SO_2

2. Is each molecule polar or nonpolar? (Reminder: you determined their shapes in the exercises for the last section.)

   a. NBr_3
   b. SF_2
   c. SiH_4

Answers

1. a. polar
   b. polar
   c. polar

2. a. polar
   b. polar
   c. nonpolar

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8: Counting Atoms, Ions, and Molecules

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8.1: Counting Atoms by the Gram

> **◐ Learning Objectives**
>
> - Define the mole and explain the need for this chemical counting unit.
> - Use Avogadro's number to convert between moles and number of particles of an element.
> - Use the molar mass to convert between grams and number of moles of an element.

When objects are very small, it is often inconvenient, inefficient, or even impossible to deal with the objects one at a time. For these reasons we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number and Mole

In chemistry, it is impossible to deal with a single atom or molecule because we can't see them, count them, or weigh them. Chemists have selected a number of particles with which to work that is convenient. Since molecules are extremely small, you may suspect this number is going to be very large, and you are right. The number of particles in this group is 6.022×10^{23} particles and the name of this group is the **mole** (the abbreviation for *mole* is mol). One mole of any object is 6.022×10^{23} of those objects. There is a particular reason that this number was chosen and this reason will become clear as we proceed.

When chemists are carrying out chemical reactions, it is important that the relationship between the numbers of particles of each reactant is known. Any readily measurable mass of an element or compound contains an extraordinarily large number of atoms, molecules, or ions, so an extremely large numerical unit is needed to count them. The mole is used for this purpose.

The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly 6.02214076×10²³ particles, e.g., atoms, molecules, ions, or electrons. The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 (^{12}C) (the isotope of carbon with relative atomic mass 12 Daltons, by definition). For most purposes, 6.022×10^{23} provides an adequate number of significant figures. Just as 1 mole of atoms contains 6.022×10^{23} atoms, 1 mole of eggs contains 6.022×10^{23} eggs. This number is called Avogadro's number.

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.022×10^{23}.

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles by using the relationship

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms}. \tag{8.1.1}$$

> **✔ Example 8.1.1: Moles of Carbon**
>
> The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?
>
> **Solution**
>
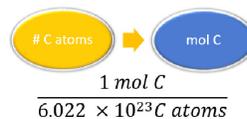> | Steps for Problem Solving | The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon? |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given: 4.72×10^{24} C atoms
Find: mol C |
> | List other known quantities. | $1\,mol = 6.022 \times 10^{23}$ C atoms |
> | | |

Steps for Problem Solving	The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?
Prepare a concept map and use the proper conversion factor.	 $$\frac{1 \; mol \; C}{6.022 \; \times 10^{23} C \; atoms}$$
Cancel units and calculate.	$$4.72 \times 10^{24} \;\; C \; atoms \times \frac{1 \; mol \; C}{6.022 \times 10^{23} \;\; C \; atoms} = 7.84 \; mol \; C$$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole.

Molar Mass

Molar mass is defined as the mass of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of the element lithium is 6.94g, the molar mass of zinc is 65.38g, and the molar mass of gold is 196.97g. Each of these quantities contains 6.022×10^{23} atoms of that particular element. The units for molar mass are grams per mole or g/mol. $1.00 \; mol$ of carbon-12 atoms has a mass of 12.0 g and contains 6.022×10^{23} atoms. 1.00 mole of any element has a mass numerically equal to its atomic mass in grams and contains 6.022×10^{23} particles. The mass, in grams, of 1 mole of particles of a substance is now called the **molar mass** (mass of 1.00 mole).

Converting Grams to Moles of an Element and Vice Versa

We can also convert back and forth between grams of an element and moles. The conversion factor for this is the molar mass of the substance. The **molar mass** is the ratio giving the number of grams for each one mole of the substance. This ratio is easily found by referring to the atomic mass of the element using the periodic table. This ratio has units of grams per mole or g/mol.

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 6.4.1 illustrates what conversion factor is needed and two examples are given below.
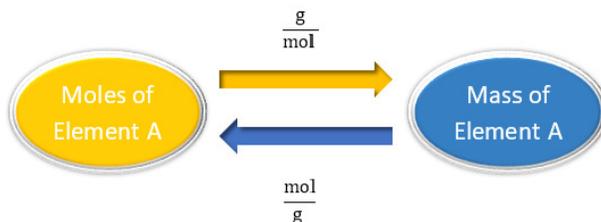
Figure 8.1.1: A Simple Flowchart for Converting Between Mass and Moles of an Element.

✔ Example 8.1.2: Chromium

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

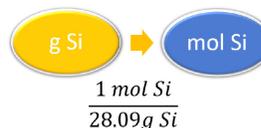| Steps for Problem Solving	Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.
Identify the "given" information and what the problem is asking you to "find."	Given: 0.560 mol Cr Find: g Cr
List other known quantities.	1 mol Cr = 52.00g Cr

Steps for Problem Solving	Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.
Prepare a concept map and use the proper conversion factor.	(mol Cr) ➡ (g Cr) $\dfrac{52.00g\ Cr}{1\ mol\ Cr}$
Cancel units and calculate.	$0.560\ \text{mol Cr} \times \dfrac{52.00\ \text{g Cr}}{1\ \text{mol Cr}} = 29.1\ \text{g Cr}$
Think about your result.	Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because of the 0.560 mol
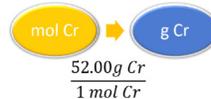
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✔ Example 8.1.3: Silicon

How many moles are in 107.6g of Si?

Solution

Steps for Problem Solving	How many moles are in 107.6 g of Si.
Identify the "given" information and what the problem is asking you to "find."	Given: 107.6 g Si Find: mol Si
List other known quantities.	1 mol Si = 28.09g Si
Prepare a concept map and use the proper conversion factor.	(g Si) ➡ (mol Si) $\dfrac{1\ mol\ Si}{28.09g\ Si}$
Cancel units and calculate.	$107.6\ \text{g Si} \times \dfrac{1\ \text{mol Si}}{28.09\ \text{g Si}} = 3.83\ \text{mol Si}$ (8.1
Think about your result.	Since 1 mol of Si is 28.09g, 107.6 should be about 4 moles.

? Exercise 8.1.1

a. How many moles are present in 100.0 g of Al?
b. What is the mass of 0.552 mol of Ag metal?

Answer a:

   3.706 mol Al

Answer b:

   59.5 g Ag

Contributions & Attributions

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- Henry Agnew (UC Davis)

- Wikipedia

8.2: Counting Molecules by the Gram

> **Learning Objectives**
> - Define molecular mass and formula mass.
> - Convert between the mass of the substance and number of moles.
> - Convert between the mass of the substance and the number of particles.

Molecular and Formula Masses

The molecular mass of a substance is the sum of the average masses of the atoms in one molecule of a substance. It is calculated by adding together the atomic masses of the elements in the substance, each multiplied by its subscript (written or implied) in the molecular formula. Because the units of atomic mass are atomic mass units, the units of molecular mass are also atomic mass units. The procedure for calculating molecular masses is illustrated in Example 8.2.1.

> **✔ Example 8.2.1: Ethanol**
>
> Calculate the molecular mass of ethanol, whose condensed structural formula is CH_3CH_2OH. Among its many uses, ethanol is a fuel for internal combustion engines
>
> **Solution**
>
> | Steps for Problem Solving | Calculate the molecular mass of ethanol, whose condensed structural formula is CH_3CH_2OH |
> |---|---|
> | Identify the "given"information and what the problem is asking you to "find." | Given: Ethanol molecule (CH₃CH₂OH)
Find: molecular mass |
> | Determine the number of atoms of each element in the molecule. | The molecular formula of ethanol may be written in three different ways:
• CH₃CH₂OH (which illustrates the presence of an ethyl group)
• CH₃CH₂–, and an –OH group)
• C₂H₅OH, and C₂H₆O;
All show that ethanol has two carbon atoms, six hydrogen atoms, and one oxygen atom. |
> | Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element. | 1 C atom = 12.011 amu
1 H atom = 1.0079 amu
1 O atom = 15.9994 amu |
> | Add the masses together to obtain the molecular mass. | 2C: (2 atoms)(12.011amu/atom) = 24.022 amu
6H: (6 atoms)(1.0079amu/atom) = 6.0474amu
+1O: (1 atoms)(15.9994amu/atom) =15.9994amu
C₂H₆O : molecular mass of ethanol = 46.069amu |

> **? Exercise 8.2.1: Freon**
>
> Calculate the molecular mass of trichlorofluoromethane, also known as Freon-11, which has a condensed structural formula of CCl_3F. Until recently, it was used as a refrigerant. The structure of a molecule of Freon-11 is as follows:
>
>
>
> Freon-11, CCl₃F
>
> Figure 8.2.1: Molecular structure of freon-11, CCl_3F.
>
> **Answer**
> 137.37 amu
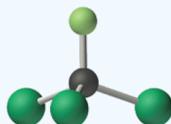
Unlike molecules, which form covalent bonds, ionic compounds do not have a readily identifiable molecular unit. Therefore, for ionic compounds, the **formula mass** (also called the empirical formula mass) of the compound is used instead of the molecular mass. The formula mass is the sum of the atomic masses of all the elements in the empirical formula, each multiplied by its subscript (written or implied). It is directly analogous to the molecular mass of a covalent compound. The units are atomic mass units.

> *Atomic mass, molecular mass, and formula mass all have the same units: atomic mass units.*

> **✔ Example 8.2.2: Calcium Phosphate**
>
> Calculate the formula mass of $Ca_3(PO_4)_2$, commonly called calcium phosphate. This compound is the principal source of calcium found in bovine milk.
>
> **Solution**
>
> | Steps for Problem Solving | Calculate the formula mass of $Ca_3(PO_4)_2$, commonly called calcium phosphate. |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given: Calcium phosphate [Ca₃(PO₄)₂] formula unit
Find: formula mass |
> | Determine the number of atoms of each element in the molecule. | • The empirical formula—Ca₃(PO₄)₂—indicates that the simplest electrically neutral unit of calcium phosphate contains three Ca²⁺ ions and two PO₄³⁻ ions.
• The formula mass of this molecular unit is calculated by adding together the atomic masses of three calcium atoms, two phosphorus atoms, and eight oxygen atoms. |
> | Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element. | 1 Ca atom = 40.078 amu
1 P atom = 30.973761 amu
1 O atom = 15.9994 amu |

Steps for Problem Solving	Calculate the formula mass of $Ca_3(PO_4)_2$, commonly called calcium phosphate.
Add together the masses to give the formula mass.	3Ca: (3 atoms) (40.078 amu/atom)=120.234amu 2P: (2 atoms) (30.973761amu/atom)=61.947522amu + 8O: (8 atoms)(15.9994amu/atom)=127.9952amu Formula mass of Ca₃(PO₄)₂=310.177amu

---

? Exercise 8.2.2: Silicon Nitride

Calculate the formula mass of Si_3N_4, commonly called silicon nitride. It is an extremely hard and inert material that is used to make cutting tools for machining hard metal alloys.

Figure 8.2.2: Si_3N_4 bearing parts. (Public Domain; David W. Richerson and Douglas W. Freitag; Oak Ridge National Laboratory).

Answer

140.29 amu

Molar Mass

The molar mass of a substance is defined as the mass in grams of 1 mole of that substance. One mole of isotopically pure carbon-12 has a mass of 12 g. For an element, the molar mass is the mass of 1 mol of atoms of that element; for a covalent molecular compound, it is the mass of 1 mol of molecules of that compound; for an ionic compound, it is the mass of 1 mol of formula units. That is, the molar mass of a substance is the mass (in grams per mole) of 6.022×10^{23} atoms, molecules, or formula units of that substance. In each case, the number of grams in 1 mol is the same as the number of atomic mass units that describe the atomic mass, the molecular mass, or the formula mass, respectively.

> *The molar mass of any substance is its atomic mass, molecular mass, or formula mass in grams per mole.*

The periodic table lists the atomic mass of carbon as 12.011 amu; the average molar mass of carbon—the mass of 6.022×10^{23} carbon atoms—is therefore 12.011 g/mol:

Table 8.2.1: Molar Mass of Select Substances

Substance (formula)	Basic Unit	Atomic, Molecular, or Formula Mass (amu)	Molar Mass (g/mol)
carbon (C)	atom	12.011 (atomic mass)	12.011
ethanol (C₂H₅OH)	molecule	46.069 (molecular mass)	46.069
calcium phosphate [Ca₃(PO₄)₂]	formula unit	310.177 (formula mass)	310.177

Converting Between Grams and Moles of a Compound

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment you need 3.00 moles of calcium chloride $(CaCl_2)$. Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of $CaCl_2$ that you should measure as shown in Example 8.2.3.

✔ Example 8.2.3: Calcium Chloride

Calculate the mass of 3.00 moles of calcium chloride (CaCl₂).

Figure 8.2.3: Calcium chloride is used as a drying agent and as a road deicer.

Solution

Steps for Problem Solving	Calculate the mass of 3.00 moles of calcium chloride (CaCl2).
Identify the "given" information and what the problem is asking you to "find."	Given: 3.00 moles of CaCl₂ Find: g CaCl₂
List other known quantities.	1 mol CaCl₂ = 110.98 g CaCl₂
Prepare a concept map and use the proper conversion factor.	$$\dfrac{110.98 \ g \ CaCl_2}{1 \ mol \ CaCl_2}$$
Cancel units and calculate.	$3.00 \ \cancel{mol \ CaCl_2} \times \dfrac{110.98 \ g \ CaCl_2}{1 \ \cancel{mol \ CaCl_2}} = 333 \ g \ CaCl_2$
Think about your result.	

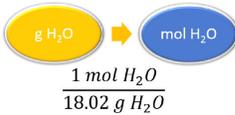**? Exercise 8.2.3: Calcium Oxide**

What is the mass of 7.50 mol of (calcium oxide) CaO?

Answer

   421 g

✔ Example 8.2.4: Water

How many moles are present in 108 grams of water?

Solution

Steps for Problem Solving	How many moles are present in 108 grams of water?
Identify the "given" information and what the problem is asking you to "find."	Given: 108 g H_2O Find: mol H_2O
List other known quantities.	1 mol H_2O = 18.02 g H_2O
Prepare a concept map and use the proper conversion factor.	 $\dfrac{1 \; mol \; H_2O}{18.02 \; g \; H_2O}$
Cancel units and calculate.	$108 \; \cancel{g \, H_2O} \times \dfrac{1 \text{ mol } H_2O}{18.02 \; \cancel{g \, H_2O}} = 5.99 \text{ mol } H_2O$
Think about your result.	

? Exercise 8.2.4: Nitrogen Gas

What is the mass of 7.50 mol of Nitrogen gas N_2?

Answer

   210. g

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to moles. To convert from mass to number of particles or vice-versa, it will first require a conversion to moles as shown in Figure $8.2.1$ and Example $8.2.5$.
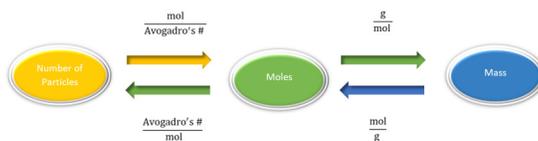
Figure 8.2.4: Conversion from number of particles to mass, or from mass to number of particles, requires two steps.

✔ Example 8.2.5: Chlorine

How many molecules is 20.0 g of chlorine gas, Cl_2?

Solution

Steps for Problem Solving	How many molecules is 20.0 g of chlorine gas, Cl_2?
Identify the "given" information and what the problem is asking you to "find."	Given: 20.0 g Cl_2 Find: # Cl_2 molecules
List other known quantities.	1 mol Cl_2 = 70.90 g Cl_2, 1mol Cl_2 = 6.022 x 10^{23} Cl_2 molecules
Prepare a concept map and use the proper conversion factors.	
Cancel units and calculate.	$20.0 \; \cancel{g \, Cl_2} \times \dfrac{1 \; \cancel{mol \, Cl_2}}{70.90 \; \cancel{g \, Cl_2}} \times \dfrac{6.022 \times 10^{23} \text{ molecules } Cl_2}{1 \; \cancel{mol \, Cl_2}} = 1.70 \times 10^{23} \text{ molec}$
Think about your result.	Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

✔ Example 8.2.6: Zinc Bromide

What is the mass of 9.376×10^{23} formula units of zinc bromide?

Solution

Steps for Problem Solving	What is the mass of 9.376×10^{23} formula units of zinc bromide?
Identify the "given" information and what the problem is asking you to "find."	Given: 9.376×10^{23} formula units of $ZnBr_2$ Find: g $ZnBr_2$
List other known quantities.	1 mol $ZnBr_2$ = 225.19 g $ZnBr_2$ 1mol $ZnBr_2$ = 6.022 x 10^{23} $ZnBr_2$ formula units
Prepare a concept map and use the proper conversion factors.	
Cancel units and calculate.	9.376×10^{23} ~~formula units $ZnBr_2$~~ $\times \dfrac{1 \text{ mol } ZnBr_2}{6.022 \times 10^{23} \text{ formula units } ZnBr_2} \times$
Think about your result.	Since the number of formula units is larger than Avogadro's number, the mass is larger than the formula mass of zinc bromide.
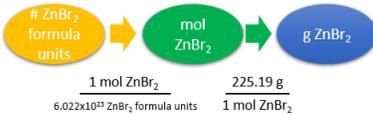
❓ Exercise 8.2.5: Calcium Chloride

How many formula units are in 25.0 g of $CaCl_2$?

Answer

    1.36 x 10^{23} $CaCl_2$ formula units

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8.3: Chemical Formulas as Conversion Factors

Learning Objectives

- Use chemical formulas as conversion factors.
- Given a compound formula and the mass of the substance calculate the mass of constituent elements and number of component particles (elements or ions).

Figure $8.3.1$ 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.
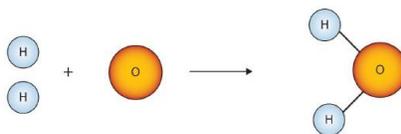
Figure 8.3.1 Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

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 $8.3.1$).

Table 8.3.1: 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	$\dfrac{2\,\text{mol H atoms}}{1\,\text{mol H}_2\text{O molecules}}$ or $\dfrac{1\,\text{mol H}_2\text{O molecules}}{2\,\text{mol H atoms}}$
1 O atom	1 mol of O atoms	$\dfrac{1\,\text{mol O atoms}}{1\,\text{mol H}_2\text{O molecules}}$ or $\dfrac{1\,\text{mol H}_2\text{O molecules}}{1\,\text{mol O atoms}}$

⚑ 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; https://what-if.xkcd.com/4/).
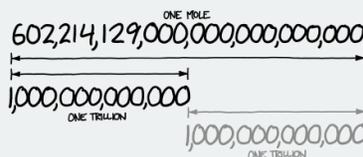
A trillion trillion kilograms is how much a planet weighs. If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Table 8.3.2: 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	$\dfrac{2\,\text{mol C atoms}}{1\,\text{mol C}_2\text{H}_6\text{O molecules}}$ or $\dfrac{1\,\text{mol C}_2\text{H}_6\text{O molecules}}{2\,\text{mol C atoms}}$
6 H atoms	6 mol of H atoms	$\dfrac{6\,\text{mol H atoms}}{1\,\text{mol C}_2\text{H}_6\text{O molecules}}$ or $\dfrac{1\,\text{mol C}_2\text{H}_6\text{O molecules}}{6\,\text{mol H atoms}}$
1 O atom	1 mol of O atoms	$\dfrac{1\,\text{mol O atoms}}{1\,\text{mol C}_2\text{H}_6\text{O molecules}}$ or $\dfrac{1\,\text{mol C}_2\text{H}_6\text{O molecules}}{1\,\text{mol O atoms}}$
2 (12.01 amu) C 24.02 amu C	2 (12.01 g) C 24.02 g C	$\dfrac{24.02\,\text{g C}}{1\,\text{mol C}_2\text{H}_6\text{O molecules}}$ or $\dfrac{1\,\text{mol C}_2\text{H}_6\text{O molecules}}{24.02\,\text{g C}}$
6 (1.008 amu) H 6.048 amu H	6 (1.008 g) H 6.048 g H	$\dfrac{6.048\,\text{g H}}{1\,\text{mol C}_2\text{H}_6\text{O molecules}}$ or $\dfrac{1\,\text{mol C}_2\text{H}_6\text{O molecules}}{6.048\,\text{g H}}$

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular and Mass Relationships	
1 (16.00 amu) O 16.00 amu O	1 (16.00 g) O 16.00 g O	$\dfrac{16.00 \text{ g O}}{1 \text{ mol } C_2H_6O \text{ molecules}}$ $\dfrac{1 \text{ mol } C_2H_6O \text{ molecules}}{16.00 \text{ g O}}$	or

The following example illustrates how we can use the relationships in Table $8.3.2$ as conversion factors.

✔ Example 8.3.1: Ethanol

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms does it have?

Solution

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.	 $\dfrac{2 \ mol \ C}{1 \ mol \ C_2H_6O}$
Cancel units and calculate.	Note how the unit *mol C_2H_6O molecules* cancels algebraically. $2.5 \ \text{mol } C_2H_6O \text{ molecules} \times \dfrac{2 \text{ mol C atoms}}{1 \text{ mol } C_2H_6O \text{ molecules}} = 5.0 \text{ mol C atoms}$
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 8.3.1

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

Once mass or moles of a compound is converted to the moles of a component an element or ion , Avogadro's number can be used to find the number of individual atoms or ions. This calculalculation, like the other ones covered in this section, can also be done in reverse.

✔ Example 8.3.2: Aluminum Oxide Mass

What mass of aluminum oxide do you need to measure out to have a sample with 4.32×10^{21} aluminum ions?

Solution

Steps for Problem Solving	What mass of aluminum oxide do you need to measure out to have a sample with 4.32×10^{21} aluminum ions?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.32×10^{21} Al^{3+} Find: g Al_2O_3
List other known quantities.	1 mol Al^{3+} = 6.022×10^{23} Al^{3+} 1 mol Al_2O_3 = 2 mol Al^{3+} 1 mol Al_2O_3 = 101.96 g Al_2O_3
Prepare a concept map and use the proper conversion factor.	 $\dfrac{1 \text{ mol } Al^{3+}}{6.022\text{x}10^{23} \ Al^{3+}}$ $\dfrac{1 \text{ mol } Al_2O_3}{2 \text{ mol } Al^{3+}}$ $\dfrac{101.96 \text{ g}}{1 \text{ mol } Al_2O_3}$
Cancel units and calculate.	$4.32 \times 10^{21} \ Al^{3+} \times \dfrac{1 \ \text{mol } Al^{3+}}{6.022 \times 10^{23} \ Al^{3+}} \times \dfrac{1 \ \text{mol } Al_2O_3}{2 \ \text{mol } Al^{3+}} \times \dfrac{101.96 \text{ g } Al_2O_3}{1 \ \text{mol } Al_2O_3} =$

? Exercise 8.3.2

What number of carbon atoms is present in 82.17 g of C_4H_{10}?

Answer:

3.406×10^{23} C

Alternatively, once you convert from mass of a compound to moles of a particular element, you can use the moles of the element to find the mass contributed to the total by that element alone. This calculation can be done in reverse as well.

✔ Example 8.3.3: Oxygen Mass

Determine the mass of Oxygen in 75.0g of C_2H_6O.

Solution

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.	
Cancel units and calculate.	$75.0 \ \text{g } C_2H_6O \times \dfrac{1 \ \text{mol } C_2H_6O}{46.07 \ \text{g } C_2H_6O} \times \dfrac{1 \ \text{mol } O}{1 \ \text{mol } C_2H_6O} \times \dfrac{16.00 \ \text{g O}}{1 \ \text{mol } O} = 26.0 \ \text{g O}$
Think about your result.	The mass of oxygen has to be less than the total mass of the ethanol.
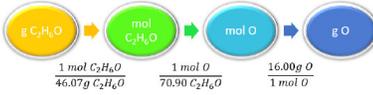
? Exercise 8.3.3

What mass of sodium sulfate contains 18.0 g of sodium?

Answer:

55.6 g Na_2SO_4

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8.4: Mass Percent Composition of Compounds

◀▶ Learning Objectives

- Determine percent composition of each element in a compound based on mass.

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter 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 that of the composition of the peanut butter.

$$\% \text{ by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\% \qquad (8.4.1)$$

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

✔ Example 8.4.1: Percent Composition from Mass Data

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 g of zinc remains. Determine the percent composition of the compound.

Solution

Steps for Problem Solving	When a 20.00 g sample of the zinc-and-oxygen compound is decomposed, 16.07 g of zinc remains. Determine the percent composition of the compound.
Identify the "given" information and what the problem is asking you to "find."	Given : Mass of compound = 20.00 g Mass of Zn = 16.07 g Find: % Composition (% Zn and %O)
List other known quantities.	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. Mass of oxygen = 20.00 g - 16.07 g = 3.93 g O
Cancel units and calculate.	$$\% \text{ Zn} = \frac{16.07 \text{ g Zn}}{20.00 \text{ g}} \times 100\% = 80.35\% \text{ Zn}$$ $$\% \text{ O} = \frac{3.93 \text{ g O}}{20.00 \text{ g}} \times 100\% = 19.65\% \text{ O}$$ Calculate the percent by mass of each element by dividing the mass of that element by the mass of the compound and multiplying by 100%.
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.

> **?** Exercise 8.4.1
>
> Sulfuric acid, H_2SO_4 is a very useful chemical in industrial processes. If 196.0 g of sulfuric acid contained 64.0 g sulfur and 4.0 g of hydrogen, what is the percent composition of the compound?
>
> **Answer**
>
> 2.0% H, 32.7% S, and 65.3% O

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8.5: Mass Percent Composition from a Chemical Formula

> ◨ **Learning Objectives**
>
> - Given a chemical formula calculate the percent composition of each component element.

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. This is divided by the molar mass of the compound and multiplied by 100%.

$$\% \text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\% \tag{8.5.1}$$

The percent composition of a given compound is always the same, given that the compound is pure.

> ✔ **Example 8.5.1**
>
> Dichlorine heptoxide $\left(\mathrm{Cl_2O_7}\right)$ is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.
>
> **Solution**
>
> | Steps for Problem Solving | Calculate the percent composition of dichlorine heptoxide $\left(\mathbf{Cl_2O_7}\right)$. |
> |---|---|
> | Identify the "given" information and what the problem is asking you to "find." | Given : $\mathrm{Cl_2O_7}$
 Find: % Composition (% Cl and %O) |
> | List other known quantities. | Mass of Cl in 1 mol $\mathrm{Cl_2O_7}$, 2 Cl : 2 x 35.45 g = 70.90 g
 Mass of O in 1 mol $\mathrm{Cl_2O_7}$, 7 O: 7 x 16.00 g = 112.00 g
 Molar mass of $\mathrm{Cl_2O_7}$ = 182.90 g/mol |
> | Cancel units and calculate. | $$\%\mathrm{Cl} = \frac{70.90 \text{ g Cl}}{182.90 \text{ g}} \times 100\% = 38.76\% \text{ Cl}$$ $$\% \mathrm{O} = \frac{112.00 \text{ g O}}{182.90 \text{ g}} \times 100\% = 61.24\% \text{ O}$$ 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%. |
> | 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% Cl and 61.24% O. Suppose that you needed 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.

$$12.50 \text{ g Cl}_2\mathrm{O}_7 \times \frac{38.76 \text{ g Cl}}{100 \text{ g Cl}_2\mathrm{O}_7} = 4.845 \text{ g Cl} \tag{8.5.2}$$

$$12.50 \text{ g Cl}_2\mathrm{O}_7 \times \frac{61.24 \text{ g O}}{100 \text{ g Cl}_2\mathrm{O}_7} = 7.655 \text{ g O} \tag{8.5.3}$$

The sum of the two masses is 12.50 g the mass of the sample size.

? Exercise 8.5.1

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a:

78.32% Ba and 21.67% F

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CHAPTER OVERVIEW

9: An Introduction to Chemical Reactions

9.1: Chemical Equations

> **◐ Learning Objectives**
>
> - Identify the reactants and products in any chemical reaction.
> - Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that—what is produced—or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\mathrm{Flour + Butter + Sugar + Vanilla + Baking\ Soda + Eggs + Chocolate\ Chips}}_{\mathrm{Ingredients = Reactants}} \rightarrow \underbrace{\mathrm{Cookies}}_{\mathrm{Product}}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\mathrm{SO_2 + O_2}$, are reactants and sulfur trioxide, $\mathrm{SO_3}$, is the product.

$$\underbrace{2\,\mathrm{SO_2(g) + O_2(g)}}_{\mathrm{Reactants}} \rightarrow \underbrace{2\,\mathrm{SO_3(g)}}_{\mathrm{Products}}$$
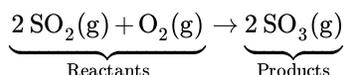
In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\mathrm{Reactants} \rightarrow \mathrm{Products}$$
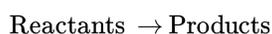
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table $9.1.1$ shows a listing of symbols used in chemical equations.

Table 9.1.1: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state
$\overset{\mathrm{Pt}}{\rightarrow}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)

Symbol	Description	Symbol	Description
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

    1. They could draw a picture of the chemical reaction.
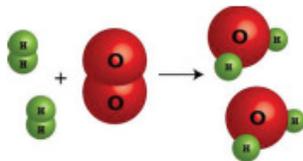
    2. They could write a word equation for the chemical reaction:

"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

    3. They could write a balanced chemical equation.

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g) \tag{9.1.1}$$

In the chemical equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the balanced chemical equation is the quickest and clearest method for representing chemical reactions.

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9.2: Balancing Chemical Equations

> **Learning Objectives**
>
> - Explain the roles of subscripts and coefficients in chemical equations.
> - Given an unbalanced chemical equation balance the 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.
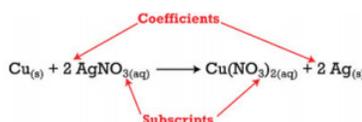
Figure 9.2.1: Balancing Equations. *You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (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 9.2.1.
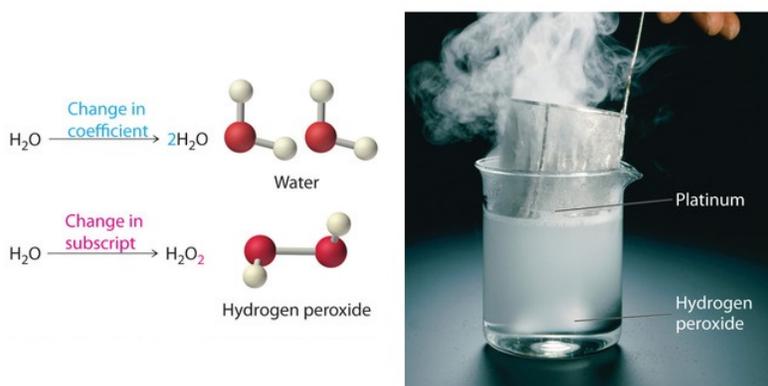
The simplest and most generally useful method for balancing chemical equations is "inspection," better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

> ⌻ **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 9.2.1: Combustion of Heptane**
>
> Balance the chemical equation for the combustion of Heptane (C_7H_{16}).
>
> $$C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$
>
> **Solution**

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 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $$C_7H_{16}(l) + O_2(g) \rightarrow \underline{7}CO_2(g) + H_2O(g)$$ • 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $$C_7H_{16}(l) + O_2(g) \rightarrow 7\,CO_2(g) + \underline{8}H_2O(g)$$ • 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: $$C_7H_{16}(l) + \underline{11}O_2(g) \rightarrow 7\,CO_2(g) + 8\,H_2O(g)$$ • 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 9.2.2: Combustion of Isooctane**

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$
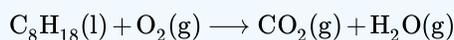
Solution

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.

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 9.2.3: 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

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $$Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$$
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. $$Pb(NO_3)_2(aq) + \underline{2}NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)$$ • 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: $$Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow \underline{2}NaNO_3(aq) + PbCl_2(s)$$ • 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.**	$$Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow 2\,NaNO_3(aq) + PbCl_2(s)$$ • 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 9.2.1

Is each chemical equation balanced?

a. $2\,Hg(\ell) + O_2(g) \rightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)$
c. $Mg(NO_3)_2(s) + 2\,Li(s) \rightarrow Mg(s) + 2\,LiNO_3(s)$

Answer a

yes

Answer b

no

Answer c

yes

? Exercise 9.2.2

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
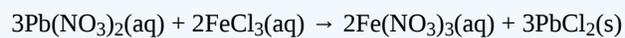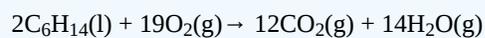
$N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$

Answer b

$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

Answer c

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

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9.3: Writing Chemical Equations

> ⬤ **Learning Objectives**
>
> - Given a sentence that describes a chemical reaction write a balanced chemical equation.

To turn a description of a chemical reaction into a balanced chemical equation, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in the chapters covering ionic compounds and covalent compounds.
3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table below.
4. Add phase labels to indicate whether the substances are solids, liquids, gases, or aqueous.
5. Finally, balance the equation. This should always be the final step.

Table 9.3.1: *Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

> ✔ **Example 9.3.1**
>
> Transfer the following symbolic equations into word equations or word equations into symbolic equations.
>
> a. Gaseous propane, C_3H_8, burns in oxygen to produce gaseous carbon dioxide and water vapor.
> b. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.
>
> **Solution**
>
> a.
>
> Reactants: propane (C_3H_8) and oxygen (O_2)
>
> Products: carbon dioxide (CO_2) and water (H_2O)
>
> Unbalanced chemical equation:
>
> $$C_3H_8\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)$$
>
> Balanced chemical equation:
>
> $$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(l)$$
>
> b.
>
> Reactants: hydrogen fluoride (HF) and potassium carbonate (K_2CO_3)
>
> Products: potassium fluoride (KF), water (H_2O), and carbon dioxide (CO_2)
>
> Unbalanced chemical equation:
>
> $$HF\,(g) + K_2CO_3\,(aq) \rightarrow KF\,(aq) + H_2O\,(l) + CO_2\,(g)$$
>
> Balanced chemical equation:
>
> $$2HF\,(g) + K_2CO_3\,(aq) \rightarrow 2KF\,(aq) + H_2O\,(l) + CO_2\,(g)$$
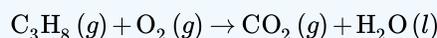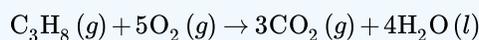
? Exercise 9.3.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

  a. Hydrogen reacts with nitrogen to produce gaseous ammonia.
  b. Copper metal is heated with oxygen to produce solid copper(II) oxide.

Answer a
$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$
Answer b
$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$
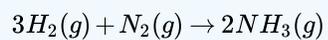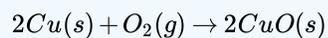
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9.4: Classifying Chemical Reactions- Take One

> **Learning Objectives**
>
> - Given a chemical equation classify the equation as a combination, decomposition, single replacement, or double replacement reaction.

The chemical reactions we have described are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Many chemical reactions can be classified into one or more of four basic types: combination, decomposition, single replacement, or double replacement. The general forms of these four kinds of reactions are summarized in Table 9.4.1, along with examples of each.

Table 9.4.1: *Basic Types of Chemical Reactions*

Name of Reaction	General Form	Examples
Single Replacement	AB + C → AC + B	$ZnCl_2(aq)$+ Mg(s) → $MgCl_2$(aq)+ Zn(s)
Double Replacement	AB + CD → AD + CB	$BaCl_2$(aq) + Na_2SO_4(aq) → $BaSO_4$(s) + 2NaCl(aq)
Combination	A + B → AB	CO_2(g) + H_2O(l) → H_2CO_3(aq) N_2(g) + $2O_2$(g)→ $2NO_2$(g)
Decomposition	AB → A + B	$CaCO_3$(s) → CaO(s) + CO_2(g)

The classification scheme is only for convenience; the same reaction can be classified in different ways, depending on which of its characteristics is most important. We will see another way of classifying reactions on the next page.

Combination Reactions

A **combination reaction** is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$$A + B \rightarrow AB \tag{9.4.1}$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

$$2Na\,(s) + Cl_2\,(g) \rightarrow 2NaCl\,(s)$$
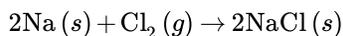
Notice that in order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2).

One type of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide:

$$2Mg\,(s) + O_2\,(g) \rightarrow 2MgO\,(s)$$
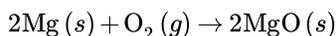
Sulfur reacts with oxygen to form sulfur dioxide:

$$S\,(s) + O_2\,(g) \rightarrow SO_2\,(g)$$
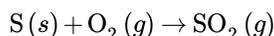
When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide:

$$2S\,(s) + 3O_2\,(g) \rightarrow 2SO_3\,(g)$$
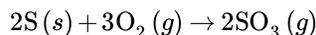
Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide:

$$2\text{Fe}\,(s) + \text{O}_2\,(g) \rightarrow 2\text{FeO}\,(s)$$

$$4\text{Fe}\,(s) + 3\text{O}_2\,(g) \rightarrow 2\text{Fe}_2\text{O}_3\,(s)$$
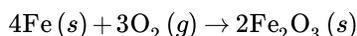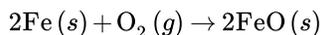
Combination reactions can also take place when an element reacts with a compound to form a new compound composed of a larger number of atoms. Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation:

$$2\text{CO}\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g) \tag{9.4.2}$$

Two compounds may also react to form a more complex compound. A very common example is the reactions of oxides with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide:

$$\text{CaO}\,(s) + \text{H}_2\text{O}\,(l) \rightarrow \text{Ca(OH)}_2\,(aq) \tag{9.4.3}$$
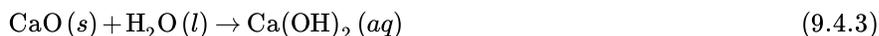
Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$\text{SO}_3\,(g) + \text{H}_2\text{O}\,(l) \rightarrow \text{H}_2\text{SO}_4\,(aq) \tag{9.4.4}$$

Figure 9.4.1: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain.

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$\text{AB} \rightarrow \text{A} + \text{B} \tag{9.4.5}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas:

$$2\text{HgO}\,(s) \rightarrow 2\text{Hg}\,(l) + \text{O}_2\,(g) \tag{9.4.6}$$

Video* 9.4.2: Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.*

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide:

$$\text{CaCO}_3\,(s) \rightarrow \text{CaO}\,(s) + \text{CO}_2\,(g) \tag{9.4.7}$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water:

$$2\text{NaOH}\,(s) \rightarrow \text{Na}_2\text{O}\,(s) + \text{H}_2\text{O}\,(g) \tag{9.4.8}$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water:

$$\text{H}_2\text{CO}_3\,(aq) \rightarrow \text{CO}_2\,(g) + \text{H}_2\text{O}\,(l) \tag{9.4.9}$$

Single Replacement Reactions

A third type of reaction is the single replacement reaction, in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$\text{A} + \text{BC} \rightarrow \text{AC} + \text{B} \tag{9.4.10}$$

In this general reaction, element A is a metal and replaces element B, also a metal, in the compound. In some cases, hydorgen can take the place of a metal. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$\text{Y} + \text{XZ} \rightarrow \text{XY} + \text{Z} \tag{9.4.11}$$

where Y is a nonmetal and replaces the nonmetal Z in the compound with X.

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\text{Mg}\,(s) + \text{Cu(NO}_3)_2\,(aq) \rightarrow \text{Mg(NO}_3)_2\,(aq) + \text{Cu}\,(s) \tag{9.4.12}$$
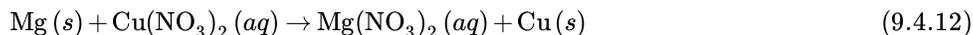
This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$\text{Zn}\,(s) + 2\text{HCl}\,(aq) \rightarrow \text{ZnCl}_2\,(aq) + \text{H}_2\,(g) \tag{9.4.13}$$
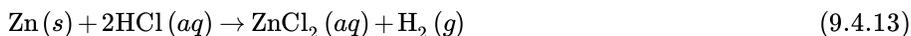
Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All Group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2\text{Na}\,(s) + 2\text{H}_2\text{O}\,(l) \rightarrow 2\text{NaOH}\,(aq) + \text{H}_2\,(g) \tag{9.4.14}$$
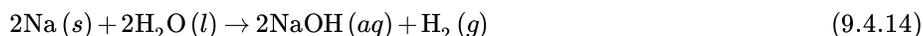
Figure 9.4.2: (First image) Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Second image) Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite.

Nonmetal Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine:

$$\text{Cl}_2\,(g) + 2\text{NaBr}\,(aq) \rightarrow 2\text{NaCl}\,(aq) + \text{Br}_2\,(l) \tag{9.4.15}$$
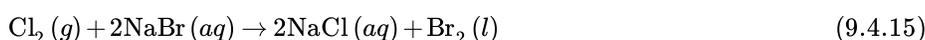
The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

Double Replacement Reactions

A **double-replacement reaction** is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement (also called double-displacement) reaction is:

$$\text{AB} + \text{CD} \rightarrow \text{AD} + \text{BC} \tag{9.4.16}$$

In this reaction, A and C are positively-charged cations, while B and D are negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

The following are all examples of double replacement reactions:
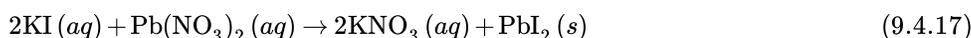
$$2\text{KI}\,(aq) + \text{Pb(NO}_3)_2\,(aq) \rightarrow 2\text{KNO}_3\,(aq) + \text{PbI}_2\,(s) \tag{9.4.17}$$

$$\text{Na}_2\text{S}\,(aq) + 2\text{HCl}\,(aq) \rightarrow 2\text{NaCl}\,(aq) + \text{H}_2\text{S}\,(g) \tag{9.4.18}$$

$$\text{HCl}\,(aq) + \text{NaOH}\,(aq) \rightarrow \text{NaCl}\,(aq) + \text{H}_2\text{O}\,(l) \tag{9.4.19}$$

9.5: Classifying Chemical Reactions- Take Two

> ◑ **Learning Objectives**
>
> - Classify a chemical reaction as a precipitation reaction, neutralization reaction, or redox reaction.

In the previous section, we classified chemical reactions according to the pattern of how the atoms rearranged. Another way to classify reactions is based on the types of products and reactants. We will examine three types of reactions in this manner: precipitation reactions, neutralization reactions, and redox reactions. A brief introduction to each type of reaction is given bellow, followed by a page looking at the type of more reaction in more depth.

Precipitation Reactions

In a precipitation reaction, the reactants are two aqueous (soluble) ionic compounds. One of the products is an insoluble ionic compound; the second product is an ionic compound that remains aqueous. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs:

$$2KI\,(aq) + Pb(NO_3)_2\,(aq) \rightarrow 2KNO_3\,(aq) + PbI_2\,(s) \tag{9.5.1}$$
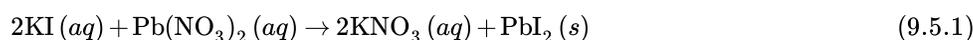
There are very strong attractive forces that occur between Pb^{2+} and I^- ions and the result is a brilliant yellow precipitate (Figure 9.5.3). The other product of the reaction, potassium nitrate, remains soluble.

Figure 9.5.3: Lead (II) iodide precipitates when potassium iodide is mixed with lead (II) nitrate (Equation 9.5.1). (CC BY-SA 3.0; PRHaney).

Here are several more examples of precipitation reactions. Notice that the pattern of two soluble ionic compounds reacting to form a insoluble ionic compound and a soluble ionic compounds is repeated in each reactions. Also notice that the order of products is arbitrary: the insoluble ionic compound can be written as the first product or as the second product.

$$BaCl_2\,(aq) + Na_2SO_4\,(aq) \rightarrow 2NaCl\,(aq) + BaSO_4\,(s) \tag{9.5.2}$$

$$Cu(NO_3)_2\,(aq) + Na_2CO_3\,(aq) \rightarrow CuCO_3\,(s) + 2NaNO_3\,(aq) \tag{9.5.3}$$

$$3NaOH\,(aq) + AlCl_3\,(aq) \rightarrow Al(OH)_3\,(s) + 3NaCl\,(aq) \tag{9.5.4}$$
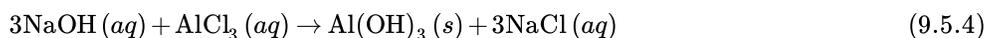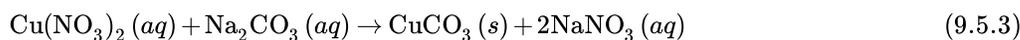
Neutralization Reactions

The reactants in a neutralization reaction are an acid and a hydroxide salt (another name for any ionic compound that contains hydroxide ions). One of the products is always water, and the other is a new ionic compound. Notice in the following examples that the hydroxide salt can be either solid or aqueous and that the order of both the reactants and the products is arbitrary.

$$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l) \tag{9.5.5}$$

$$2KOH\,(s) + H_2SO_4\,(aq) \rightarrow K_2SO_4\,(aq) + 2H_2O\,(l) \tag{9.5.6}$$

$$Ba(OH)_2\,(aq) + 2HNO_3\,(aq) \rightarrow 2H_2O\,(l) + BaSO_4\,(aq) \tag{9.5.7}$$
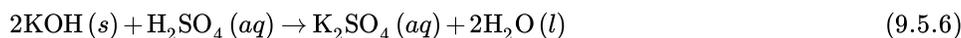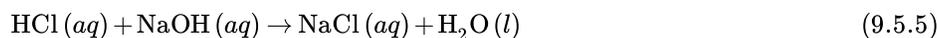
Redox Reactions (Oxidation-Reduction Reactions)

To simplify discussion of this complex topic, we will only look at redox reactions in which an atom appears in its elemental form on one side of a reaction and as part of a compound on the other side of the reaction. The reaction of heptane and oxygen gives water vapor and carbon dioxide:

$$\text{C}_7\text{H}_{16}\,(l) + 11\text{O}_2\,(g) \rightarrow 7\text{CO}_2\,(g) + 8\text{H}_2\text{O}\,(g) \tag{9.5.8}$$
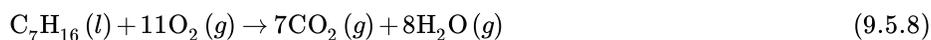
To identify this as a redox reaction, focus on the oxygen atoms. On the reactant side, they appear in the form of elemental oxygen. On the product side, they appear in two compounds: carbon dioxide and water.

In the examples that follow notice that the elements may be monoatomic or diatomic They may appear as products or reactants. There may also be two or more types of atoms transitioning from their elemental form to part of a compound (or vice versa). Also, you do not need to pay attention to the phase labels when identifying a redox reaction.

$$2\text{Na}\,(s) + \text{Cl}_2\,(g) \rightarrow 2\text{NaCl}\,(s) \tag{9.5.9}$$

$$2\text{CO}\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g) \tag{9.5.10}$$

$$2\text{HgO}\,(s) \rightarrow 2\text{Hg}\,(l) + \text{O}_2\,(g) \tag{9.5.11}$$

$$\text{Zn}\,(s) + 2\text{HCl}\,(aq) \rightarrow \text{ZnCl}_2\,(aq) + \text{H}_2\,(g) \tag{9.5.12}$$
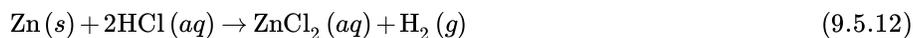
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9.6: Solubility Rules

> **◑ Learning Objectives**
>
> - Given a table of solubility rules and the name or formula for an ionic compound, determine whether the ionic compound is soluble in water.

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (Tables 9.6.1 and 9.6.2).

Table 9.6.1: Solubility Rules for Soluble Substances

Soluble in Water	Important Exceptions (Insoluble)
All Group IA and NH_4^+ salts	none
All nitrates, chlorates, perchlorates and acetates	none
All sulfates	$CaSO_4$, $BaSO_4$, $SrSO_4$, $PbSO_4$
All chlorides, bromides, and iodides	AgX, Hg_2X_2, PbX_2 (X= Cl, Br, or I)
All fluorides	MgF_2, CaF_2, SrF_2, BaF_2, PbF_2,

Table 9.6.2: Solubility Rules for Sparingly Soluble Substances

Sparingly Soluble in Water	Important Exceptions (Soluble)
All carbonates, phosphates, and chromates	Group IA and NH_4^+ salts
All oxides and hydroxides	Group IA and NH_4^+ salts; Ba^{2+}, Sr^{2+}, Ca^{2+} sparingly soluble
All sulfides	Group IA, IIA and NH_4^+ salts; MgS, CaS, BaS sparingly soluble
All oxalates	Group IA and NH_4^+ salts

As an example on how to use the solubility rules, let's predict if cesium nitrate and lead(II) bromide are soluble in water. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, lead(II) bromide is insoluble in water.

> **✔ Example 9.6.1: Solubility**
>
> Classify each compound as soluble or insoluble
>
> a. $Zn(NO_3)_2$
> b. $PbBr_2$
> c. $Sr_3(PO_4)_2$
>
> **Solution**
>
> a. All nitrates are soluble in water, so $Zn(NO_3)_2$ is soluble.
> b. All bromides are soluble in water, except those combined with Pb^{2+}, so $PbBr_2$ is insoluble.
> c. All phosphates are insoluble, so $Sr_3(PO_4)_2$ is insoluble.

? Exercise 9.6.1: Solubility

Classify each compound as soluble or insoluble.

a. Mg(OH)$_2$
b. KBr
c. Pb(NO$_3$)$_2$

Answer a

insoluble

Answer b

soluble

Answer c

soluble

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9.7: Precipitation Reactions

> **◐ Learning Objectives**
>
> - Given the reactants in an aqueous chemical reaction, predict the products in a double displacement reaction.

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.

$$\text{AgNO}_3(\text{aq}) + \text{K}_2\text{Cr}_2\text{O}_7(\text{aq}) \rightarrow \text{Ag}_2\text{Cr}_2\text{O}_7(\text{s}) + \text{KNO}_3(\text{aq}) \tag{9.7.1}$$
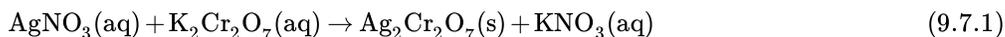
This unbalanced equation has the general form of an exchange reaction:

$$\text{AC} + \text{BD} \rightarrow \underset{insoluble}{\text{AD}} + \text{BC} \tag{9.7.2}$$

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 9.7.1).

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 $\text{Na}^+(\text{aq})$., $\text{Cl}^-(\text{aq})$., $\text{K}^+(\text{aq})$., and $\text{Br}^-(\text{aq})$. As you will see in (Figure 9.7.1), 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.
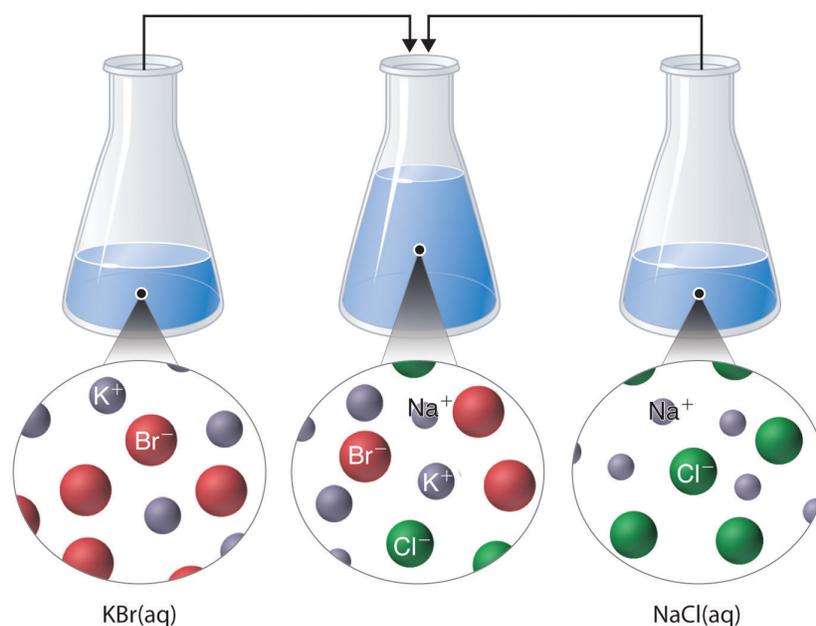
Figure 9.7.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

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):
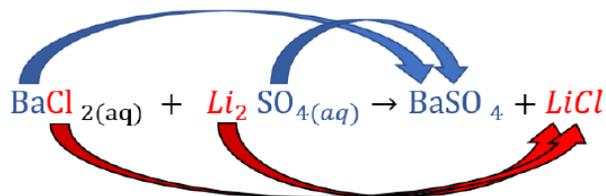
Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains Ba^{2+}, Cl^-, Li^+, and SO_4^{2-} ions. The only possible exchange reaction is to form LiCl and $BaSO_4$.

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.
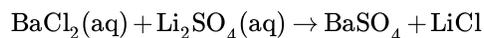
$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4 + LiCl$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + LiCl(aq)$$
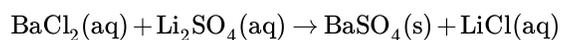
The solubility rules from the previous section shows that LiCl is soluble in water, but $\ce{BaSO_4}$ is not soluble in water.

Balance the equation:

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,LiCl(aq)$$
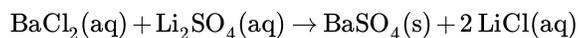
Although soluble barium salts are toxic, $BaSO_4$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a

"barium milkshake" or a "barium enema"—a suspension of very fine $BaSO_4$ particles in water.

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 BaSO4 particles in water; the high atomic mass of barium makes it opaque to x-rays. from Wikipedia.

✔ Example 9.7.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

Solution

Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	$Rb\,OH_{(aq)} + CoCl_{2(aq)} \rightarrow RbCl + Co(OH)_2$
Correct the formulas of the products based on the charges of the ions.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2(aq) + Co(OH)_2(s)$
Balance the equation.	Coefficients already balanced. $Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$

✔ Example 9.7.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution
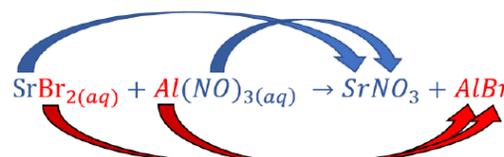
Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	$SrBr_{2(aq)} + Al(NO)_{3(aq)} \rightarrow SrNO_3 + AlBr$
Correct the formulas of the products based on the charges of the ions.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2 + AlBr_3$

Steps	Example
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2(aq) + AlBr_3(aq)$ According to the solubility rules from the previous section, both AlBr$_3$ (rule 4) and Sr(NO$_3$)$_2$ (rule 2) are soluble.
If all possible products are soluble, then no net reaction will occur.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow$ **NO REACTION**

? Exercise 9.7.2

Using the information in the solubility rules 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.

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9.8: Complete Ionic and Net Ionic Equations- Precipitation Reaction Examples

> **◗◖ Learning Objectives**
>
> - Represent a precipitation reaction using a complete ionic or net ionic equation.

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations.

$$\underbrace{\mathrm{BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2\,NaCl(aq)}}_{\text{Complete Chemical Equation}} \tag{9.8.1}$$
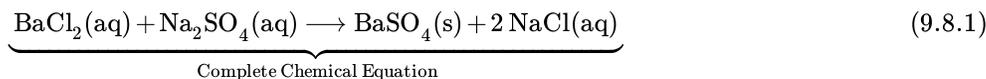
However, Equation 9.8.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Below is the complete ionic equation:

$$\underbrace{\mathrm{Ba^{2+}(aq) + \overbrace{2\,Cl^-(aq)}^{\text{spectator}} + \overbrace{2\,Na^+(aq)}^{\text{spectator}} + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + \overbrace{2\,Na^+(aq)}^{\text{spectator}} + \overbrace{Cl^-(aq)}^{\text{spectator}}}}_{\text{Complete Ionic Equation}} \tag{9.8.2}$$
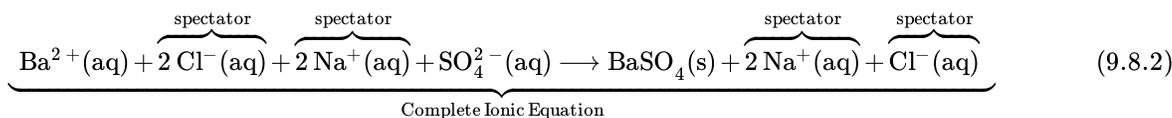
Equation 9.8.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $\mathrm{BaSO_4(s)}$ will form when *any* solution containing $\mathrm{Ba^{2+}(aq)}$ is mixed with *any* solution containing $\mathrm{SO_4^{2-}(aq)}$ (provided concentrations are not extremely small). This happens independently of the $\mathrm{Cl^-(aq)}$ and $\mathrm{Na^+(aq)}$ ions in Equation 9.8.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $\mathrm{BaSO_4}$ the net ionic equation is

$$\underbrace{\mathrm{Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)}}_{\text{Net Ionic Equation}} \tag{9.8.3}$$
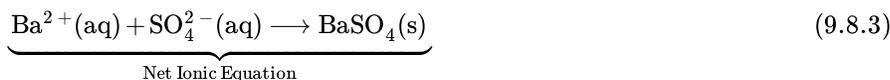
> **✔ Example 9.8.1**
>
> a. When a solution of $\mathrm{AgNO_3}$ is added to a solution of $\mathrm{CaCl_2}$, insoluble AgCl precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.
> b. Write the balanced net ionic equation to describe any reaction that occurs when the solutions of $\mathrm{Na_2SO_4}$ and $\mathrm{NH_4I}$ are mixed.
>
> **Solution**
>
> | Equation Type | Example 9.8.1a | Example 9.8.1b |
> |---|---|---|
> | Complete Chemical Equation | $\mathrm{2\,AgNO_3(aq) + CaCl_2(aq) \longrightarrow}$ $\mathrm{2\,AgCl(s) + Ca(NO_3)_2(aq)}$
 The proper states and formulas of all products are written and the chemical equation is balanced. | $\mathrm{Na_2SO_4(aq) + NH_4I_2(aq) \longrightarrow}$ $\mathrm{2\,NaI(aq) + (NH_4)_2SO_4(aq)}$
 Both products are aqueous so there is no net ionic equation that can be written. |
> | Complete Ionic Equation | $\mathrm{2\,Ag^+(aq) + 2\,NO_3^-(aq) + Ca^{2+}(aq) + Cl^-(aq) \longrightarrow}$ $\mathrm{2\,AgCl(s) + Ca^{2+}(aq) + 2\,NO_3^-(aq)}$
 AgCl is a solid so it does not break up into ions in solution. | |
> | Net Ionic Equation | $\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$
 All spectator ions are removed. | NaI and $\mathrm{(NH_4)_2SO_4}$ are both soluble.
 No net ionic equation |

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $BaCl_2$ solution, for instance, is often used as a test for the presence of $SO_4^{2-}(aq)$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $BaSO_4$. Thus, if $BaCl_2$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the SO_4^{2-} ion.

Figure 9.8.1: The three common silver halide precipitates: AgI, AgBr and AgCl (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$AgNO_3$ solutions are often used in a similar way to test for halide ions. If $AgNO_3$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of Cl^- ions, a cream-colored precipitate indicates the presence of Br^- ions, and a yellow precipitate indicates the presence of I^- ions (Figure 9.8.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $AgNO_3$ is added to tap water, a white precipitate is almost always formed. The Cl^- ions in tap water usually come from the Cl_2 which is added to municipal water supplies to kill microorganisms.

> **? Exercise 9.8.1**
>
> Write balanced net ionic equations to describe any reaction that occurs when the following solutions are mixed.
>
> a. $K_2CO_3 + SrCl_2$
> b. $FeSO_4 + Ba(NO_3)_2$
>
> **Answer a**
>
> $$Sr^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow SrCO_3(s)$$
>
> **Answer b**
>
> $$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Ba(SO_4)(s)$$
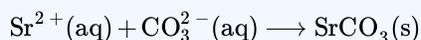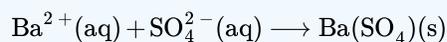
Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then the rules of stoichiometry may be applied.

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9.8: Complete Ionic and Net Ionic Equations- Precipitation Reaction Examples is shared under a not declared license and was authored, remixed, and/or curated by LibreTexts.

9.9: Complete Ionic and Net Ionic Equations - More Examples

> ◑ **Learning Objectives**
>
> - Convert a formula (molecular) equation into a full ionic or net ionic equation.

When converting a formula (molecular) equation into an ionic equation, remember the following guidelines:

- Always start with a balanced formula (molecular) equation.
- Only soluble ionic compounds dissociate into ions. We will deal with acids--the only significant exception to this guideline--next term.
- When disassociating an ionic compound into its component ions, be carefuly not pull apart polyatomic ions. For example, if a compound contains nitrate ions, don't convert NO_3^-(aq) to N^{3-}(aq) and $3O^{2-}$(aq). Neither nitride ions nor oxide ions are stable in an aqueous environment. Also, examination of the charges should convince you that nitrate is not composed of oxide and nitride ions.
- Be careful with ionic compounds that have multiple monatomic ions in a formula unit. For example, $CaCl_2$ dissociates into one calcium ion and two chloride ions: Ca^{2+}(aq) + $2Cl^-$(aq). Don't be tempted to write Cl_2^-(aq). The formula Cl_2^- would indicate a polyatomic ion formed from two chlorine atoms that has an overall charge of -1; it does not indicate two chloride ions.

It might be helpful to look at a few more examples.

> ✔ **Example 9.9.1**
>
> In the following section, we will examine the reaction that occurs when a solid piece of elemental magnesium in placed in an aqueous solution of copper(II) chloride:
>
> $$CuCl_2(aq) + Mg(s) \rightarrow Cu(s) + MgCl_2(aq) \tag{9.9.1}$$
>
> Write the full ionic and net ionic equations for this reaction.
>
> **Solution**
>
> The elemental metals (magnesium on the reactant side, copper on the product side) are neutral solids. They therefore appear unaltered in the full ionic equation. Only the aqueous ionic compounds (the two chloride salts) are written as ions:
>
> $$Cu^{2+}(aq) + 2\,Cl^-(aq) + Mg(s) \rightarrow Cu(s) + Mg^{2+}(aq) + 2\,Cl^-(aq) \tag{9.9.2}$$
>
> The chloride ions are spectator ions. To get the net ionic equation, we cancel them from both sides of the equation:
>
> $$Cu^{2+}(aq) + Mg(s) \rightarrow Cu(s) + Mg^{2+}(aq) \tag{9.9.3}$$
>
> Note that when variable-charge metals such as copper appear as part of a compound, we have to determine the charge on the cation by looking at the number of anions and their charge. This is the same process we followed when naming a compound with a variable-charge metal in chapter 4.
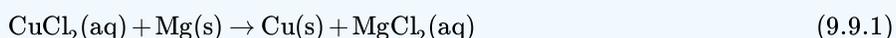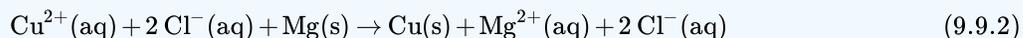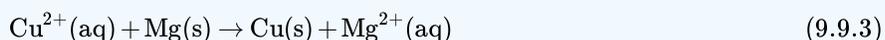
> ✔ **Example 9.9.2**
>
> When an excess of an aqueous hydroxide salt is added to a solution containing ammonium ions, ammonia gas is formed:
>
> $$2\,NH_4Cl(aq) + Ba(OH)_2(aq) \rightarrow 2\,NH_3(g) + BaCl_2(aq) + 2\,H_2O(l) \tag{9.9.4}$$
>
> Write the full ionic and net ionic equations for this reaction.
>
> **Solution**
>
> Both the compounds on the reactant side of the equation are soluble ionic compounds, so they will need to be separated into their respective ions. In this case, both compounds contain a polyatomic ion. Remember, these polyatomic ions maintain their integrity in solution; do not separate them into ions. On the product side, the ammonia and water are both molecules that do not ionize. Only the barium chloride is separated into ions:
>
> $$2\,NH_4^+(aq) + 2\,Cl^-(aq) + Ba^{2+}(aq) + 2\,OH^-(aq) \rightarrow 2\,NH_3(g) + Ba^{2+}(aq) + 2\,Cl^-(aq) + 2\,H_2O(l) \tag{9.9.5}$$
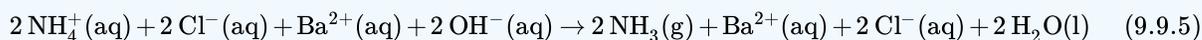
Both the barium ions and the chloride ions are spectator ions. The net ionic equation results from cancelling them from the full ionic equation:

$$2\,NH_4^+\,(aq) + 2\,OH^-\,(aq) \rightarrow 2\,NH_3\,(g) + 2\,H_2O\,(l) \tag{9.9.6}$$
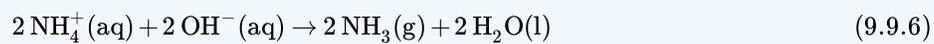
In the case of this net ionic equation, the stoicheometric coefficients can be reduced by dividing through by two:

$$NH_4^+\,(aq) + OH^-\,(aq) \rightarrow NH_3\,(g) + H_2O\,(l) \tag{9.9.7}$$
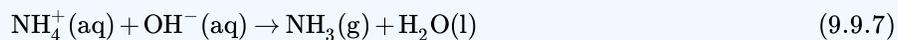
The mechanism of the reaction becomes more clear by inspecting the net ionic equation: the ammonia molecule is created from the ammonium ion when the hydroxide ion strips a hydrogen away from it. By gaining a hydrogen (and a unit of charge) the hydroxide ion transforms into a water molecule.

9.10: A Very Brief Introduction to Redox Reactions

> **Learning Objectives**
>
> - Associate the term redox with electron transfer.
> - Recognize that a redox reaction is occurring when an atom is present in its elemental state on one side of a reaction and as part of a compound on the other side of the reaction.

"Redox" is short for "oxidation and reduction." Redox reactions are a large family of chemical reactions that involve transfer of electrons from one species to another. To simplify this complex topic for now, we will only look at redox reactions when an atom appears in its elemental form on one side of a chemical reaction and as part of a compound on the other side.

In all oxidation-reduction reactions an exchange of electrons occurs - one substance loses electrons while something else gains them. That is the key to understanding redox reactions.

Figure 9.10.1: *Reaction of copper wire in a silver nitrate solution.*

A simple demonstration of a redox reaction involves placing a solid piece of copper wire in a silver nitrate solution. Within minutes the wire begins to look fuzzy or furry, as small silver crystals begin to form on the wire. Meanwhile, the originally clear silver nitrate solution begins to take on a pale bluish tint. Furthermore, if the crystals are shaken off of the wire we see that the wire partially disintegrated.

The overall equation for our demonstration describes the events:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,Ag(s) \tag{9.10.1}$$
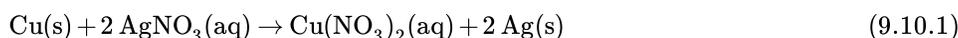
In a simplified representation of this reaction, we can look at what changes are occuring with the copper and silver atoms:

$$Cu(s) + 2\,Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2\,Ag(s) \tag{9.10.2}$$
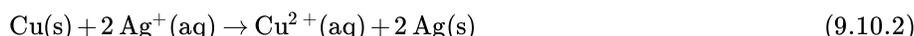
The copper is changing from its neutral, elemental form to one of its ionic forms. For the silver, it is changing from its ionic form to its elemental form. These transformations necessitate an exchange of electrons.

Oxidation of Copper Metal to Make Copper Ions

Copper began as a neutral atom with no charge, but changed into an ion with a charge of +2. An atom becomes a positive ion by losing electrons:

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2\,e^- \tag{9.10.3}$$

Notice that copper began as a solid but is converted into aqueous ions - this is why the copper wire disintegrates. We say that copper was **oxidized** because is has lost electrons (i.e., electrons appear on the product side of the Equation 9.10.3).

Reduction of Silver Ions to Make Silver Metal

Silver was converted from an ion with a charge of +1, Ag+, to a neutral atom, Ag. The only way an ion can undergo this change is to gain an electron:

$$Ag^+(aq) + e^- \rightarrow Ag(s) \tag{9.10.4}$$

Notice that solid silver is formed - this is what causes the fuzzy appearance to begin appearing on the wire - solid silver crystals. Silver has gained electrons - it has been reduced (i.e., electrons appear on the reactant side of Equation 9.10.4).

The electrons that silver gained had to come from somewhere - they came from copper. Conversely, a substance such as copper can only lose electrons if there is something else that will take them up, the silver ions. **One cannot occur without the other**. This exchange of electrons is what defines an **oxidation - reduction reaction**.

Oxidation is defined as the loss of electrons and **reduction** is defined as the gain of electrons.

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10: Mass Relations in Chemical Reactions

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10.1: Stoichiometry

> **◐ Learning Objectives**
>
> - Determine the relative amounts of each substance in a balanced chemical equation.

You have learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this section you will explore the quantitative relationships that exist between the quantities of reactants and products in a balanced equation. This is known as **stoichiometry**.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry is actually Greek from two words: $\sigma\tau o\iota\kappa\eta\iota o\nu$, which means "element", and \(\mu \epsilon \tau \rho \omicron \nu), which means "measure".

Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

$$2CuSO_4 + 4KI \rightarrow 2CuI + 2K_2SO_4 + I_2 \tag{10.1.1}$$
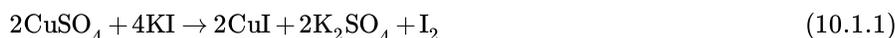
The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 2 formula units of copper (II) sulfate ($CuSO_4$) we have, we need to have 4 formula units of potassium iodide (KI). For every two dozen copper (II) sulfates, we need 4 dozen potassium iodides. Because the unit "mole" is also a counting unit, we can interpret this equation in terms of moles, as well: For every two moles of copper (II) sulfate, we need 4 moles potassium iodide.

The production of ammonia $\left(NH_3\right)$ from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$N_2\left(g\right) + 3H_2\left(g\right) \rightarrow 2NH_3\left(g\right) \tag{10.1.2}$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

1 molecule N_2 3 molecules H_2 2 molecules NH_3
1 mol N_2 3 mol H_2 2 mol NH_3
28.02 g N_2 3 x 2.02 g = 6.06 g H_2 2 x 17.04 g = 34.08 g NH_3
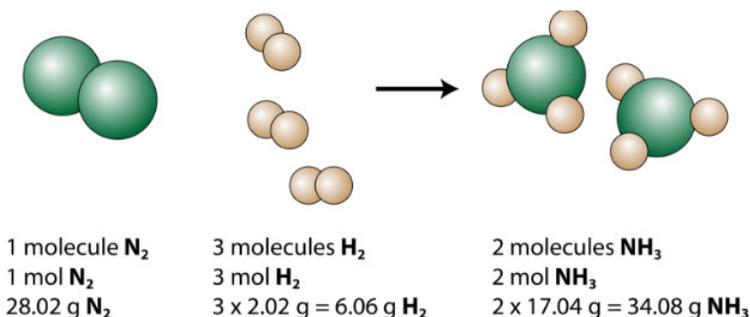
Figure 10.1.1: This representation of the production of ammonia from nitrogen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

We see that 1 molecule of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia. This is the smallest possible relative amount of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

The most useful quantity for counting particles is the mole. So if each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation.

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. 1 mol of nitrogen has a mass of 28.02 g while 3 mol of hydrogen has a mass of 6.06 g, and 2 mol of ammonia has a mass of 34.08 g

$$28.02 \text{ g N}_2 + 6.06 \text{ g H}_2 \rightarrow 34.08 \text{ g NH}_3 \qquad (10.1.3)$$
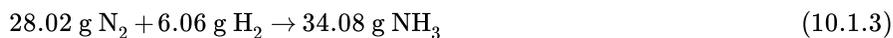
Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.

> ✔ **Example 10.1.1**
>
> The equation for the combustion of ethane (C_2H_6) is
>
> $$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$
>
> a. Indicate the number of formula units or molecules in the balanced equation.
> b. Indicate the number of moles present in the balanced equation.
>
> **Solution**
>
> a. Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.
> b. Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.
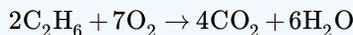
> ? **Exercise 10.1.1**
>
> For the following equation below, indicate the number of formula units or molecules, and the number of moles present in the balanced equation.
>
> $$KBrO_3 + 6KI + 6HBr \rightarrow 7KBr + 3H_2O$$
>
> **Answer**
>
> One formula unit of $KBrO_3$ plus six formula units of KI plus six molecules of HBr yields seven formula units of KBr plus three molecules of I_2 and three molecules of H_2O. One mole of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 plus three moles of H_2O.

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10.2: Mole-to-Mole Conversions

◑ Learning Objectives

- Use the mole ratios from a balanced chemical equation to solve mole to mole stoichiometry calculations.

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). As follows, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation:

$$2H_2 + O_2 \rightarrow 2H_2O \tag{10.2.1}$$

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4H_2 + 2O_2 \rightarrow 4H_2O \tag{10.2.2}$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22H_2 + 11O_2 \rightarrow 22H_2O \tag{10.2.3}$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O$$
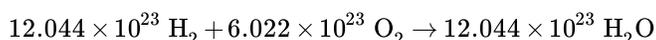
These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ mol\ H_2 + 1\ mol\ O_2 \rightarrow 2\ mol\ H_2O$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$
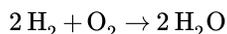
Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level, but also in terms of molar amounts of reactants and products.* Thus, we can read this reaction as "two moles of hydrogen react with one mole of oxygen to produce two moles of water."

By the same token, the ratios we constructed to describe a molecular reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2} \tag{10.2.4}$$

$$\frac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} \tag{10.2.5}$$

$$\frac{2\ \text{mol}\ H_2}{2\ \text{mol}\ H_2O}\ \text{or}\ \frac{2\ \text{mol}\ H_2O}{2\ \text{mol}\ H_2} \tag{10.2.6}$$

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called *stoichiometry*.

✔ **Example 10.2.1**

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?

Solution

Steps for Problem Solving	How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?
Find a balanced equation that describes the reaction.	Unbalanced: $H_2 + O_2 \rightarrow H_2O$ Balanced: $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$
Identify the "given" information and what the problem is asking you to "find."	Given: moles H_2O Find: moles oxygen
List other known quantities.	1 mol O_2 = 2 mol H_2O
Prepare a concept map and use the proper conversion factor.	$$\frac{1\ mol\ O_2}{2\ mol\ H_2O}$$
Cancel units and calculate.	$$27.6\ \cancel{mol\ H_2O} \times \frac{1\ mol\ O_2}{2\ \cancel{mol\ H_2O}} = 13.8\ mol\ O_2$$ To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

✔ **Example 10.2.2**

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Find a balanced equation that describes the reaction.	Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Balanced: $N_2 + \underline{3}H_2 \rightarrow \underline{2}NH_3$
Identify the "given" information and what the problem is asking you to "find."	Given: $H_2 = 4.20$ mol Find: mol of NH_3
List other known quantities.	3 mol H_2 = 2 mol NH_3
Prepare a concept map and use the proper conversion factor.	$$\frac{2\ mol\ NH_3}{3\ mol\ H_2}$$

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Cancel units and calculate.	$$4.20 \ \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3 \ \cancel{\text{mol } H_2}} = 2.80 \text{ mol } NH_3$$ The reaction of 4.20 mol of hydrogen with excess nitrogen produces 2.80 mol of ammonia.
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

? Exercise 10.2.3

a. Given the following balanced chemical equation:

$$C_5H_{12} + 8\,O_2 \rightarrow 5\,CO_2 + 6\,H_2O$$
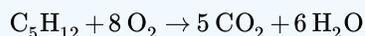
, How many moles of H_2O can be formed if 0.0652 mol of C_5H_{12} were to react?

b. Balance the following unbalanced equation and determine how many moles of H_2O are produced when 1.65 mol of NH_3 react:

$$NH_3 + O_2 \rightarrow N_2 + H_2O$$
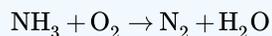
Answer a

0.391 mol H_2O

Answer b

$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol H_2O

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10.3: Mole-to-Mass, Mass-to-Mole, and Mass-to-Mass Conversions

> **◖** Learning Objectives
>
> - Use the mole ratios from a balanced chemical equation to solve mass to mass stoichiometry calculations.

Mole to Mass Conversions

We have established that a balanced chemical equation is balanced in terms of moles, as well as atoms or molecules. We have used balanced equations to set up ratios, in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions—such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that relation to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:
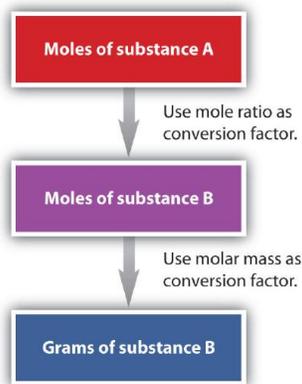
Collectively, these conversions are called mole-mass calculations.

As an example, consider the balanced chemical equation

$$Fe_2O_3 + 3SO_3 \rightarrow Fe_2(SO_4)_3 \tag{10.3.1}$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then, using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

As usual, we start with the quantity we were given:

$$3.59 \ \text{mol Fe}_2\text{O}_3 \times \left(\frac{3 \ \text{mol SO}_3}{1 \ \text{mol Fe}_2\text{O}_3} \right) = 10.77 \ \text{mol SO}_3 \tag{10.3.2}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77 \ \text{mol SO}_3 \times \left(\frac{80.06 \ \text{g SO}_3}{1 \ \text{mol SO}_3} \right) = 862 \ \text{g SO}_3 \tag{10.3.3}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

> The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:
>
> $$3.59 \ mol \ Fe_2O_3 \times \underbrace{\left(\frac{3 \ mol \ SO_3}{1 \ mol \ Fe_2O_3} \right)}_{\text{converts to moles of SO}_3} \times \underbrace{\left(\frac{80.06 \ g \ SO_3}{1 \ mol \ SO_3} \right)}_{\text{converts to grams of SO}_3} = 862 \ g \ SO_3 \tag{10.3.4}$$
>
> We get exactly the same answer when combining all math steps together.

> **✔** Example 10.3.1: Generation of Aluminum Oxide
>
> How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?
>
> $$2\,AlCl_3 + 3\,H_2O(\ell) \rightarrow Al_2O_3 + 6\,HCl(g)$$
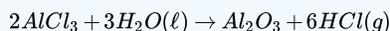
Solution

Steps for Problem Solving	Example 10.3.1
Identify the "given" information and what the problem is asking you to "find."	Given: 249 g $AlCl_3$ Find: moles HCl
List other known quantities.	1 mol $AlCl_3$ = 133.33 g/mol 6 mol of HCl to 2 mol $AlCl_3$
Prepare a concept map and use the proper conversion factor.	 $\dfrac{1\ mol\ AlCl_3}{133.33\ g\ AlCl_3}$　$\dfrac{6\ mol\ HCl}{2\ mol\ AlCl_3}$
Cancel units and calculate.	$249\ g\ \cancel{AlCl_3} \times \dfrac{1\ mol\ \cancel{AlCl_3}}{133.33\ g\ \cancel{AlCl_3}} \times \dfrac{6\ mol\ HCl}{2\ mol\ \cancel{AlCl_3}} = 5.60\ mol\ HCl$
Think about your result.	Since 249 g of $AlCl_3$ is less than 266.66 g, the mass for 2 moles of $AlCl_3$ and the relationship is 6 mol of HCl to 2 mol $AlCl_3$, the answer should be less than 6 moles of HCl.
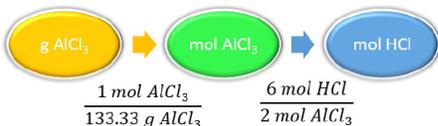
? Exercise 10.3.1: Generation of Aluminum Oxide

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$
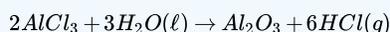
Answer

0.442 mol Al_2O_3

Mass to Mass Conversions

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:
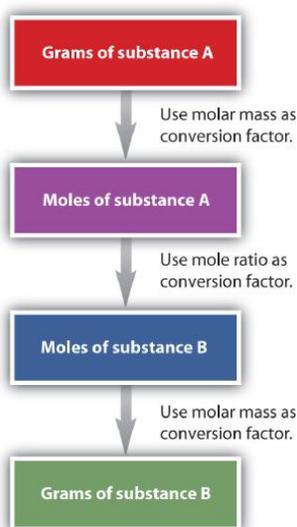
This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

✔ **Example 10.3.2: Decomposition of Ammonium Nitrate**

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$NH_4NO_3\,(s) \rightarrow N_2O\,(g) + 2H_2O\,(l)$$
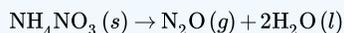
In a certain experiment, $45.7\,g$ of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Steps for Problem Solving	Example 10.3.2
Identify the "given" information and what the problem is asking you to "find."	Given: 45.7 g NH_4NO_3 Find: Mass $N_2O =?$ g Mass $H_2O =?$ g
List other known quantities.	1 mol $NH_4NO_3 = 80.06$ g/mol 1 mol $N_2O = 44.02$ g/mol 1 mol $H_2O = 18.02$ g/mol 1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O
Prepare two concept maps and use the proper conversion factor.	
Cancel units and calculate.	$45.7\text{ g NH}_4NO_3 \times \dfrac{1\text{ mol NH}_4NO_3}{80.06\text{ g NH}_4NO_3} \times \dfrac{1\text{ mol N}_2O}{1\text{ mol NH}_4NO_3} \times \dfrac{44.02\text{ g N}_2O}{1\text{ mol N}_2O} = 2$ $45.7\text{ g NH}_4NO_3 \times \dfrac{1\text{ mol NH}_4NO_3}{80.06\text{ g NH}_4NO_3} \times \dfrac{2\text{ H}_2O}{1\text{ mol NH}_4NO_3} \times \dfrac{18.02\text{ g H}_2O}{1\text{ mol H}_2O} = 2$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

? **Exercise 10.3.2: Carbon Tetrachloride**

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows:

$$CH_4(g) + 4\,Cl_2(g) \rightarrow CCl_2(l) + 4\,HCl(l) \tag{10.3.5}$$

How many grams of HCl are produced by the reaction of 100.0g of CH_4?

Answer

909.2 g HCl

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CHAPTER OVERVIEW

11: Energy and Chemical Reactions

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11.1: Energy

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we are awake or asleep.

Ability to Do Work or Produce Heat

When we speak of *using* energy, we are really referring to *transferring* energy from one place to another. When you use energy to throw a ball, you transfer energy from your body to the ball, and this causes the ball to fly through the air. When you use energy to warm your house, you transfer energy from the furnace to the air in your home, and this causes the temperature in your house to rise. Although energy is used in many kinds of different situations, all of these uses rely on energy being transferred in one of two ways. Energy can be transferred as *heat* or as *work*.

When scientists speak of **heat**, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature, as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot—the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure 11.1.1).

Figure 11.1.1: Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the food.

Heat is only one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is *force (any push or pull) applied over a distance*. When you push an object and cause it to move, you do work, and you transfer some of *your* energy to the object. At this point, it's important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but it is not true in science. By definition, scientific work requires that force be applied *over a distance*. It does not matter how hard you push or how hard you pull. If you have not moved the object, you haven't done any work.

So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work. But the question still remains—*what IS energy*?

Kinetic Energy

Machines use energy, our bodies use energy, energy comes from the sun, energy comes from volcanoes, energy causes forest fires, and energy helps us to grow food. With all of these seemingly different types of energy, it's hard to believe that there are really only

two different *forms* of energy: kinetic energy and potential energy. **Kinetic energy** is energy associated with motion. When an object is moving, it has kinetic energy. When the object stops moving, it has no kinetic energy. While all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you are from California, you might have driven through the Tehachapi Pass near Mojave or the Montezuma Hills in Solano County and seen the windmills lining the slopes of the mountains (Figure 11.1.2). These are two of the larger wind farms in North America. As wind rushes along the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.

Figure 11.1.2: A wind farm in Solano County harnesses the kinetic energy of the wind. (CC BY-SA 3.0 Unported; BDS2006 at Wikipedia)

Potential Energy

Potential energy is *stored* energy. It is energy that remains available until we choose to use it. Think of a battery in a flashlight. If left on, the flashlight battery will run out of energy within a couple of hours, and the flashlight will die. If, however, you only use the flashlight when you need it, and turn it off when you don't, the battery will last for days or even months. The battery contains a certain amount of energy, and it will power the flashlight for a certain amount of time, but because the battery stores *potential* energy, you can choose to use the energy all at once, or you can save it and only use a small amount at a time.

Any stored energy is potential energy. There are a lot of different ways in which energy can be stored, and this can make potential energy very difficult to recognize. In general, an object has potential energy because of its *position relative to another object*. For example, when a rock is held above the earth, it has potential energy because of its position relative to the ground. This is *potential energy* because the energy is *stored* for as long as the rock is held in the air. Once the rock is dropped, though, the stored energy is released as kinetic energy as the rock falls.

Chemical Energy

There are other common examples of potential energy. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When two magnets are held next to one another, they store potential energy too. For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

As you learned earlier, the world, and all of the chemicals in it are made up of atoms and molecules. These store potential energy that is dependent on their positions relative to one another. Of course, you can't see atoms and molecules. Nevertheless, scientists do know a lot about the ways in which atoms and molecules interact, and this allows them to figure out how much potential energy is stored in a specific quantity (like a cup or a gallon) of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

> Since different chemicals have different amounts of potential energy, scientists will sometimes say that potential energy depends not only on *position*, but also on *composition*. Composition affects potential energy because it determines which molecules and atoms end up next to one another. For example, the total potential energy in a cup of pure water is different than

> the total potential energy in a cup of apple juice, because the cup of water and the cup of apple juice are *composed* of different amounts of different chemicals.

At this point, you may wonder just how useful chemical potential energy is. If you want to release the potential energy stored in an object held above the ground, you just drop it. But how do you get potential energy out of chemicals? It's actually not difficult. Use the fact that different chemicals have *different amounts of potential energy*. If you start with chemicals that have a lot of potential energy and allow them to react and form chemicals with less potential energy, all the extra energy that was in the chemicals at the beginning, but not at the end, is released. The metabolism of food by our bodies, the burning of fuels such as gasoline or natural gas, and the discharging of a battery are all examples of this process.

Units of Energy

Energy is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by $1^{\circ}C$. For example, raising the temperature of $100\,g$ of water from $20^{\circ}C$ to $22^{\circ}C$ would require $100 \times 2 = 200\,cal$.

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 $85,000\,cal$ In order to make the distinction, the dietary calorie is written with a capital C.

$$1\,kilocalorie = 1\,Calorie = 1000\,calories \tag{11.1.1}$$

To say that the snack "contains" 85 Calories means that $85\,kcal$ of energy are released when that snack is processed by your 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.

$$1\,J = 0.2390\,cal\;or\;1\,cal = 4.184\,J \tag{11.1.2}$$

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

$$400\,Cal = 400\,kcal \times \frac{4.184\,kJ}{1\,kcal} = 1.67 \times 10^3\,kJ \tag{11.1.3}$$

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11.2: Temperature Changes - Specific Heat

> ◆ **Learning Objectives**
>
> - Define 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 on both 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 in the same sun will not become nearly as hot. We would say that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by $1^{\circ}C$). Water is very resistant to changes in temperature, while metals in general 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 $1^{\circ}C$. 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 $(J/g^{\circ}C)$ or calories per gram per degree $(cal/g^{\circ}C)$ (Table 11.2.1). This text will use $J/g^{\circ}C$ for specific heat.

$$\text{specific heat} = \frac{\text{heat}}{\text{mass} \times \text{temperature change}} \tag{11.2.1}$$

Notice that water has a very high specific heat compared to most other substances.

Table 11.2.1: Specific Heat Capacities

Substance	Specific Heat Capacity at 25ºC in J/g ºC	Substance	Specific Heat Capacity at 25ºC in J/g ºC
H_2 gas	14.267	steam @ 100ºC	2.010
He gas	5.300	vegetable oil	2.000
$H_2O(l)$	4.184	sodium	1.23
lithium	3.56	air	1.020
ethyl alcohol	2.460	magnesium	1.020
ethylene glycol	2.200	aluminum	0.900
ice @ 0ºC	2.010	concrete	0.880
steam @ 100ºC	2.010	glass	0.840

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 11.2.1: 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.

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11.3: Energy and Specific Heat Calculations

> **◑ Learning Objectives**
>
> - Given the mass and change in temperature of a substance and the specific heat of that substance, students will be able to use the physical property of specific heat to determine the total heat absorbed or released when a substance is heated or cooled.

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.

$$q = c_p \times m \times \Delta T \tag{11.3.1}$$

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 ΔT. 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 11.3.1 lists the specific heats for various materials.

Table 11.3.1: *Specific Heats of Some Common Substances*

Substance	Specific Heat $(\text{J/g}^\circ\text{C})$
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 = $mc\Delta T$. If energy goes into an object, the total energy of the object increases, and the values of heat ΔT are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and ΔT are negative.

✔ **Example 11.3.1**

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

Solution

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

Known

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

Unknown

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

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

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^\circ\text{C}} = 0.231 \text{ J/g}^\circ\text{C} \tag{11.3.2}$$

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 11.3.1), 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:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813 \text{ J}}{4.18 \text{ J/g}^\circ\text{C} \times 60.0 \text{ g}} = 3.24^\circ\text{C} \tag{11.3.3}$$

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

$$T_f = 23.52^\circ\text{C} - 3.24^\circ\text{C} = 20.28^\circ\text{C} \tag{11.3.4}$$

✔ **Example 11.3.2**

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:

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

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:

$$\text{heat} = (150.0 \text{ g}) \left(0.108 \; \frac{\text{cal}}{\text{g} \cdot °\text{C}} \right) (48.3°\text{C}) = 782 \text{ cal} \tag{11.3.5}$$

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 11.3.1

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 11.3.2

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 11.3.1?

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:

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

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:

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

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

$$c = 0.0923 \text{ cal/g•°C}$$

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

? Exercise 11.3.2

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

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11.4: Endothermic and Exothermic Reactions

> **◐ Learning Objectives**
>
> - Given a chemical equation with the amount of heat released written as a product students will be able to label this as an exothermic reaction.
> - Given a chemical equation with the amount of heat absorbed written as a reactant students will be able to label this as an exothermic reaction.
> - Given a chemical equation with the amount of heat released or absorbed students will be able to perform stoichiometric calculations using the molar masses and mole ratios of the substances to determine the actual amount (in Joules or calories) of heat released.

Endothermic and Exothermic Reactions

Endothermic and exothermic reactions can be thought of as having energy as either a reactant or a product. **Endothermic** reactions require energy, so energy is a reactant. Heat flows from the surroundings to the system (reaction mixture) and the energy of the system increases. In an **exothermic** reaction, heat is released (considered a product) and the energy of the system decreases.

In the course of an **endothermic** process, the system gains heat from the surroundings and so the **temperature of the surroundings decreases (gets cold).** A chemical reaction is **exothermic** if heat is released by the system into the surroundings. Because the surroundings is gaining heat from the system, **the temperature of the surroundings increases.**

Endothermic Reaction: 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. Because the heat is absorbed by the system, the 177.8 kJ is written as a **reactant**.

Heat is a reactant

$$CaCO_3 \ (s) + 177.8 \ kJ \rightarrow CaO \ (s) + CO_2 \ (g)$$
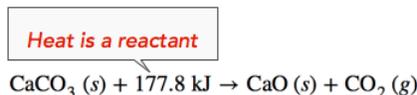
Exothermic Reaction: When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of 1 mol of methane releases 890.4 kilojoules of heat energy. The amount of heat released is therefore written on the **product** side of the reaction.

Heat is a product

$$CH_4 \ (g) + 2O_2 \ (g) \rightarrow CO_2 \ (g) + 2H_2O \ (l) + 890.4 \ kJ$$
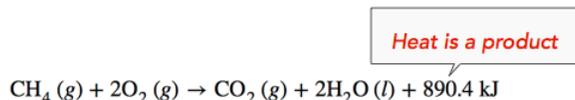
> **✔ Example 11.4.1**
>
> Is each chemical reaction exothermic or endothermic?
>
> a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 213 \ kcal$
> b. $N_2(g) + O_2(g) + 45 \ kcal \rightarrow 2NO(g)$
>
> **Solution**
>
> a. Because energy (213 kcal) is a product, energy is given off by the reaction. Therefore, this reaction is exothermic.
> b. Because energy (45 kcal) is a reactant, energy is absorbed by the reaction. Therefore, this reaction is endothermic.

> **? Exercise 11.4.2**
>
> Is each chemical reaction exothermic or endothermic?
>
> a. $H_2(g) + F_2(g) \rightarrow 2HF \ (g) + 130 \ kcal$
> b. $2C(s) + H_2(g) + 5.3 \ kcal \rightarrow C_2H_2(g)$
>
> **Answer**
>
> a. The energy (130 kcal) is produced, hence the reaction is exothermic

b. The energy (5.3 kcal) is supplied or absorbed to react, hence, the reaction is endothermic

Energy Diagrams

Endothermic and exothermic reactions can be visually represented by *energy-level diagrams* like the ones in **Figure 11.4.2**. In endothermic reactions, the **energy of the reactants is lower than that of the products**. This type of reaction is represented by an "uphill" energy-level diagram shown in Figure 11.4.2A. For an endothermic chemical reaction to proceed, the reactants must absorb energy from their environment to be converted to products.

In an exothermic reaction, the **energy of the products is lower than the energy of the reactants,** hence is energetically downhill, shown in Figure 11.4.2B. Energy is given off as reactants are converted to products. The energy given off is usually in the form of heat (although a few reactions give off energy as light). In the course of an exothermic reaction, heat flows from the system to its surroundings, and thus, gets warm.
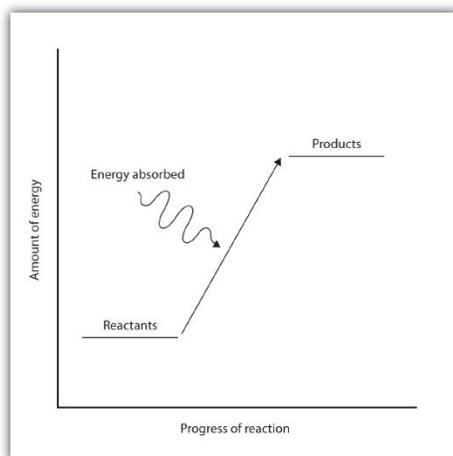
Figure 11.4.2A: Endothermic Reactions   Figure 11.4.2B: Exothermic Reactions
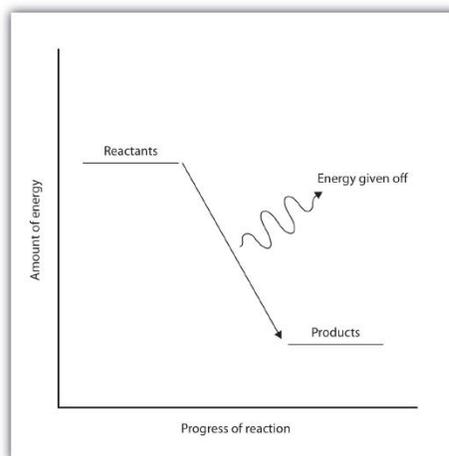
Table 11.4.2 Endothermic and Exothermic Reactions

Endothermic Reactions	Exothermic Reactions
Heat is absorbed by reactants to form products.	Heat is released.
Heat is absorbed from the surroundings; as a result, the surroundings get cold.	Heat is released by the reaction to surroundings; surroundings feel hot.
The reactants are lower in energy than the products	The products are lower in energy than the reactants
Represented by an "uphill" energy diagram	Represented by an "downhill" energy diagram

Thermochemical Stoicheometry

It is also possible to combine our previous studies of stoicheometry with thermochemical equations in order to relate the amount of energy released or absorbed during a reaction to the mass of product reacted or he mass of reactant created. The following two examples illustrate this.

> ✔ **Example 11.4.2**
>
> How much energy is needed to react 72.3 g of carbon in the presence of excess hydrogen? Acetylene, C_2H_2, is produced as the product.
>
> $2C(s) + H_2(g) + 5.3 \text{ kcal} \rightarrow C_2H_2(g)$
>
> **Solution**

If we treat the 5.3 kcal that appears in the chemical equation as another reactant, we can use it to form a stoichiometric ratios. This ratio can then be used in a dimentional analysis problem to find the solution. Don't forget to include the stoicheometric coefficient of 2 that appears in front of the carbon.

$$72.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{5.3 \text{ kcal}}{2 \text{ mol C}} = 16.0 \text{ kcal}$$

31.9 kcal of energy is therefore needed to react the 72.3 g of carbon.

✔ Example 11.4.3

Nitrogen monoxide is a pollutant produced as a byproduct whenever oxygen and nitrogen are heated together at a very high temperature. What mass of nitrogen monoxide is created when 27.51 kcal of energy is added to a mixture of oxygen and nitrogen?

$N_2(g) + O_2(g) + 45 \text{ kcal} \rightarrow 2NO(g)$

Solution

$$27.51 \text{ kcal} \times \frac{2 \text{ mol NO}}{45 \text{ kcal}} \times \frac{30.01 \text{ g NO}}{1 \text{ mol NO}} = 36.69 \text{ g NO}$$

The formation of carbon monoxide is generally an undesired side reaction. One purpose of the catalytic converter in cars and trucks is to get rid of some of the NO created in the engine as a byproduct of hydrocarbon combustion.

Exercises

1. Is each chemical reaction exothermic or endothermic?

    a. $2SnCl_2(s) + 33 \text{ kcal} \rightarrow Sn(s) + SnCl_4(s)$
    b. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 213 \text{ kcal}$

2. Is each chemical reaction exothermic or endothermic?

    a. $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g) + 137 \text{ kJ}$
    b. $C(s, \text{graphite}) + 1.9 \text{ kJ} \rightarrow C(s, \text{diamond})$

Answers

1. a: endothermic, b: exothermic

2. a: exothermic, b: endothermic

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