MAP: INTRODUCTORY CHEMISTRY (CORWIN)

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CHAPTER OVERVIEW

1: Introduction to Chemistry

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

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- **Homework**

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1.1: Evolution of Chemistry

If you are reading these words, you are likely starting a chemistry course. Get ready for a fantastic journey through a world of wonder, delight, and knowledge. One of the themes of this book is "chemistry is everywhere," and indeed it is; you would not be alive if it were not for chemistry, because your body is a big chemical machine.

Figure 1.1.1 © Thinkstock

6 images showing a volcanic eruption, large dinner spread, explosion, red wine, a pool of large fish, and molecules are laid out together.

If you do not believe it, do not worry. Every chapter in this book contains examples that will show you how chemistry is, in fact, everywhere. So enjoy the ride, and enjoy chemistry.

1.2: Modern Chemistry

Learning Objective

- Learn the basic terms used to describe matter

The definition of chemistry—the study of the interactions of matter with other matter and with energy—uses some terms that should also be defined. We start the study of chemistry by defining basic terms.

Matter

Matter is anything that has mass and takes up space. A book is matter, a computer is matter, food is matter, and dirt in the ground is matter. Sometimes matter may be difficult to identify. For example, air is matter, but because it is so thin compared to other matter (e.g., a book, a computer, food, and dirt), we sometimes forget that air has mass and takes up space. Things that are not matter include thoughts, ideas, emotions, and hopes.

Example 1.2.1

Which of the following is matter and not matter?

a. a hot dog
b. love
c. a tree

Solution

a. A hot dog has mass and takes up space, so it is matter.
b. Love is an emotion, and emotions are not matter.
c. A tree has mass and takes up space, so it is matter.

Exercise 1.2.1

Which of the following is matter and not matter?

a. the moon
b. an idea for a new invention

Answer a

The moon is matter.

Answer b

The invention itself may be matter, but the idea for it is not.

To understand matter and how it changes, we need to be able to describe matter. There are two basic ways to describe matter: physical properties and chemical properties.

Physical properties

Physical properties are characteristics that describe matter as it exists. Some physical characteristics of matter are shape, color, size, and temperature. An important physical property is the **phase** (or **state**) of matter. The three fundamental phases of matter are solid, liquid, and gas (Figure 1.2.1).

Figure 1.2.1: The Phases of Matter. Chemistry recognizes three fundamental phases of matter: solid (left), liquid (middle), and gas (right). (CC BY-SA 3.0; Spirit469). The solid state depicted is an ice sculpture of an angel, the liquid state is a drop of water, and the gas is clouds made of water vapor.

Chemical Properties

Chemical properties are characteristics of matter that describe how matter changes form in the presence of other matter. Does a sample of matter burn? Burning is a chemical property. Does it behave violently when put in water? This reaction is a chemical property as well (Figure 1.2.2). In the following chapters, we will see how descriptions of physical and chemical properties are important aspects of chemistry.

Figure 1.2.2: Chemical Properties. The fact that this match burns is a chemical property of the match. (Sebastian Ritter (Rise0011)).

Physical Change

A physical change occurs when a sample of matter changes one or more of its physical properties. For example, a solid may melt (Figure 1.2.3), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter.

Figure 1.2.2: Physical Changes: The solid ice melts into liquid water—a physical change. A time-lapse animation of ice cubes melting in a glass over 50 minutes. (Public Domain; Moussa).

Chemical Change

Chemical change is the process of demonstrating a chemical property, such as the burning match in Figure 1.2.2 "Chemical Properties". As the matter in the match burns, its chemical composition changes, and new forms of matter with new physical properties are created. Note that chemical changes are frequently accompanied by physical changes, as the new matter will likely have different physical properties from the original matter.

Example 1.2.2

Describe each process as a physical change or a chemical change.

a. Water in the air turns into snow.
b. A person's hair is cut.
c. Bread dough becomes fresh bread in an oven.

Solution

a. Because the water is going from a gas phase to a solid phase, this is a physical change.
b. Your long hair is being shortened. This is a physical change.
c. Because of the oven's temperature, chemical changes are occurring in the bread dough to make fresh bread. These are chemical changes. (In fact, a lot of cooking involves chemical changes.)

Exercise 1.2.2

Identify each process as a physical change or a chemical change.

a. A fire is raging in a fireplace.
b. Water is warmed to make a cup of coffee.

Answer a

chemical change

Answer b

physical change

Substance

A sample of matter that has the same physical and chemical properties throughout is called a substance. Sometimes the phrase *pure substance* is used, but the word *pure* isn't needed. The definition of the term *substance* is an example of how chemistry has a specific definition for a word that is used in everyday language with a different, vaguer definition. Here, we will use the term *substance* with its strict chemical definition.

> *Chemistry recognizes two different types of substances: elements and compounds.*

Element

An element is the simplest type of chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are 118 elements known to science, of which 80 are stable. (The other elements are radioactive, a condition we will consider in Chapter 15.) Each element has its own unique set of physical and chemical properties. Examples of elements include iron, carbon, and gold.

Compound

A compound is a combination of more than one element. The physical and chemical properties of a compound are different from the physical and chemical properties of its constituent elements; that is, it behaves as a completely different substance. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Mixtures

Physical combinations of more than one substance are called mixtures. Elements and compounds are not the only ways in which matter can be present. We frequently encounter objects that are physical combinations of more than one element or compound—mixtures. There are two types of mixtures.

Heterogeneous Mixture

A heterogeneous mixture is a mixture composed of two or more substances. It is easy to tell, sometimes by the naked eye, that more than one substance is present.

Homogeneous Mixture/ Solution

A homogeneous mixture is a combination of two or more substances that is so intimately mixed, that the mixture behaves as a single substance. Another word for a homogeneous mixture is a solution. Thus, a combination of salt and steel wool is a heterogeneous mixture because it is easy to see which particles of the matter are salt crystals and which are steel wool. On the other hand, if you take salt crystals and dissolve them in water, it is very difficult to tell that you have more than one substance present just by looking—even if you use a powerful microscope. The salt dissolved in water is a homogeneous mixture, or a solution (Figure 1.2.3).

Figure 1.2.3: Types of Mixtures © Thinkstock. On the left, the combination of two substances is a heterogeneous mixture because the particles of the two components look different. On the right, the salt crystals have dissolved in the water so finely that you cannot tell that salt is present. The homogeneous mixture appears like a single substance.

Example 1.2.3

Identify the following combinations as heterogeneous mixtures or homogenous mixtures.

a. soda water (carbon dioxide is dissolved in water)
b. a mixture of iron metal filings and sulfur powder (both iron and sulfur are elements)

Solution

a. Because carbon dioxide is dissolved in water, we can infer from the behavior of salt crystals dissolved in water that carbon dioxide dissolved in water is (also) a homogeneous mixture.
b. Assuming that the iron and sulfur are simply mixed together, it should be easy to see what is iron and what is sulfur, so this is a heterogeneous mixture.

Exercise 1.2.3

a. the human body
b. an amalgam, a combination of some other metals dissolved in a small amount of mercury

Answer a

heterogeneous mixture

Answer b

homogeneous mixture

There are other descriptors that we can use to describe matter, especially elements. We can usually divide elements into metals and nonmetals, and each set shares certain (but not always all) properties.

Metal

A metal is an element that conducts electricity and heat well and is shiny, silvery, solid, ductile, and malleable. At room temperature, metals are solid (although mercury is a well-known exception). A metal is ductile because it can be drawn into thin wires (a property called *ductility*); and malleable because it can be pounded into thin sheets (a property called *malleability*).

Nonmetal

A non-metal is an element that is brittle when solid, and does not conduct electricity or heat very well. Non-metals cannot be made into thin sheets or wires (Figure 1.2.4). Nonmetals also exist in a variety of phases and colors at room temperature.

Semi-metals

Some elements have properties of both metals and nonmetals and are called semi-metals (or metalloids). We will see later how these descriptions can be assigned rather easily to various elements.

Figure 1.2.4: Semimetals © Thinkstock. On the left is some elemental mercury, the only metal that exists as a liquid at room temperature. It has all the other expected properties of a metal. On the right, elemental sulfur is a yellow nonmetal that usually is found as a powder.

Describing Matter Flowchart

"Describing Matter" is a flowchart of the relationships among the different ways of describing matter.

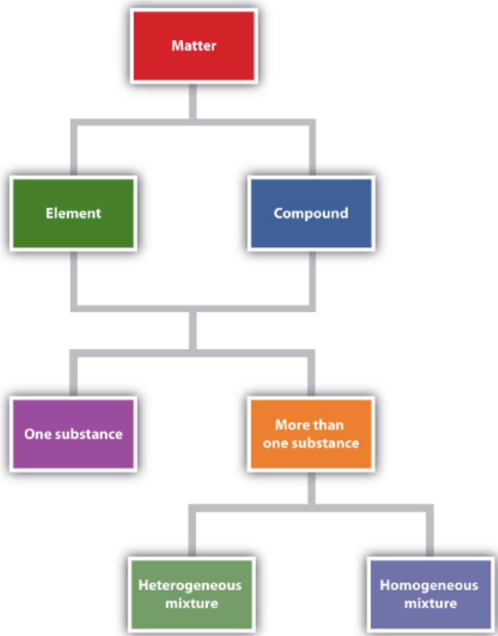


Figure 1.2.5: Describing Matter. This flowchart shows how matter can be described. Matter forks into element & compound which fork into 1 substance & more than 1 substance. More than 1 substance is homogenous or heterogenous.

Example 1.2.1: Chemistry is Everywhere: In the Morning

Most people have a morning ritual, a process that they go through every morning to get ready for the day. Chemistry appears in many of these activities.

- If you take a shower or bath in the morning, you probably use soap, shampoo, or both. These items contain chemicals that interact with the oil and dirt on your body and hair to remove them and wash them away. Many of these products also contain chemicals that make you smell good; they are called *fragrances*.
- When you brush your teeth in the morning, you usually use toothpaste, a form of soap, to clean your teeth. Toothpastes typically contain tiny, hard particles called *abrasives* that physically scrub your teeth. Many toothpastes also contain fluoride, a substance that chemically interacts with the surface of the teeth to help prevent cavities.
- Perhaps you take vitamins, supplements, or medicines every morning. Vitamins and other supplements contain chemicals your body needs in small amounts to function properly. Medicines are chemicals that help combat diseases and promote health.
- Perhaps you make some fried eggs for breakfast. Frying eggs involves heating them enough so that a chemical reaction occurs to cook the eggs.
- After you eat, the food in your stomach is chemically reacted so that the body (mostly the intestines) can absorb food, water, and other nutrients.
- If you drive or take the bus to school or work, you are using a vehicle that probably burns gasoline, a material that burns fairly easily and provides energy to power the vehicle. Recall that burning is a chemical change.

These are just a few examples of how chemistry impacts your everyday life. And we haven't even made it to lunch yet!

Figure 1.2.6: Chemistry in Real Life © Thinkstock. Examples of chemistry can be found everywhere—in personal hygiene products, food, and motor vehicles. Personal hygiene products (left), food (middle), and motor vehicles (right) images next to one another.

Key Takeaways

- Chemistry is the study of matter and its interactions with other matter and energy.
- Matter is anything that has mass and takes up space.
- Matter can be described in terms of physical properties and chemical properties.
- Physical properties and chemical properties of matter can change.
- Matter is composed of elements and compounds.
- Combinations of different substances are called mixtures.
- Elements can be described as metals, nonmetals, and semi-metals.

1.3: Learning Chemistry

Learning Objective

- Learn what science is and how it works.

Chemistry is a branch of science. Although science itself is difficult to define exactly, the following definition can serve as a starting point. Science is the process of knowing about the natural universe through observation and experiment. Science is not the only process of knowing (e.g., the ancient Greeks simply sat and *thought*), but it has evolved over more than 350 years into the best process that humanity has devised, to date, to learn about the universe around us.

The process of science is usually stated as the *scientific method*, which is rather naively described as follows:

1. state a hypothesis,
2. test the hypothesis, and
3. refine the hypothesis

In actuality, the process is not that simple. (For example, a scientist does not go into their lab every day and exclaim, "I am going to state a hypothesis today and spend the day testing it!") The process is not that simple because science and scientists have a body of knowledge that has already been identified as coming from the highest level of understanding, and most scientists build from that body of knowledge.

An educated guess about how the natural universe works is called a **hypothesis**. A scientist who is familiar with how part of the natural universe works—say, a chemist—is interested in furthering that knowledge. That person makes a reasonable guess—a hypothesis—that is designed to see if the universe works in a new way as well. Here's an example of a hypothesis: "if I mix one part of hydrogen with one part of oxygen, I can make a substance that contains both elements."

> *For a hypothesis to be termed a scientific hypothesis, it has to be something that can be supported or refuted through carefully crafted experimentation or observation.*

Most good hypotheses are grounded in previously understood knowledge and represent a testable extension of that knowledge. The scientist then devises ways to test if that hypothesis is correct or not. That is, the scientist plans experiments. **Experiments are tests of the natural universe to see if a guess (hypothesis) is correct**. An experiment to test our previous hypothesis would be to actually mix hydrogen and oxygen and see what happens. Most experiments include observations of small, well-defined parts of the natural universe designed to see results of the experiments.

A Scientific Hypothesis

A hypothesis is often written in the form of an if/then statement that gives a possibility (if) and explains what may happen because of the possibility (then). For example, if eating elemental sulfur repels ticks, then someone that is eating sulfur every day will not get ticks.

Why do we have to do experiments? Why do we have to test? Because the natural universe is not always so obvious, experiments are necessary. For example, it is fairly obvious that if you drop an object from a height, it will fall. Several hundred years ago (coincidentally, near the inception of modern science), the concept of gravity explained that test. However, is it obvious that the entire natural universe is composed of only about 115 fundamental chemical building blocks called elements? This wouldn't seem true if you looked at the world around you and saw all the different forms matter can take. In fact, the concept of *the element* is only about 200 years old, and the last naturally occurring element was identified about 80 years ago. It took decades of tests and millions of experiments to establish what the elements actually are. These are just two examples; a myriad of such examples exists in chemistry and science in general.

When enough evidence has been collected to establish a general principle of how the natural universe works, the evidence is summarized in a theory. A theory is a general statement that explains a large number of observations. "All matter is composed of atoms" is a general statement, a theory, that explains many observations in chemistry. A theory is a very powerful statement in science. There are many statements referred to as "the theory of _______" or the "______ theory" in science (where the blanks represent a word or concept). When written in this way, theories indicate that science has an overwhelming amount of evidence of its correctness. We will see several theories in the course of this text.

A specific statement that is thought to never be violated by the entire natural universe is called a law. A scientific law is the highest understanding of the natural universe that science has and is thought to be inviolate. The fact that all matter attracts all other matter—the law of gravitation—is one such law. Note that the terms *theory* and *law* used in science have slightly different meanings from those in common usage; where theory is often used to mean hypothesis ("I have a theory…"), and a law is an arbitrary limitation that can be broken but with potential consequences (such as speed limits). Here again, science uses these terms differently, and it is important to apply their proper definitions when you use these words in science. (Figure 1.3.1)

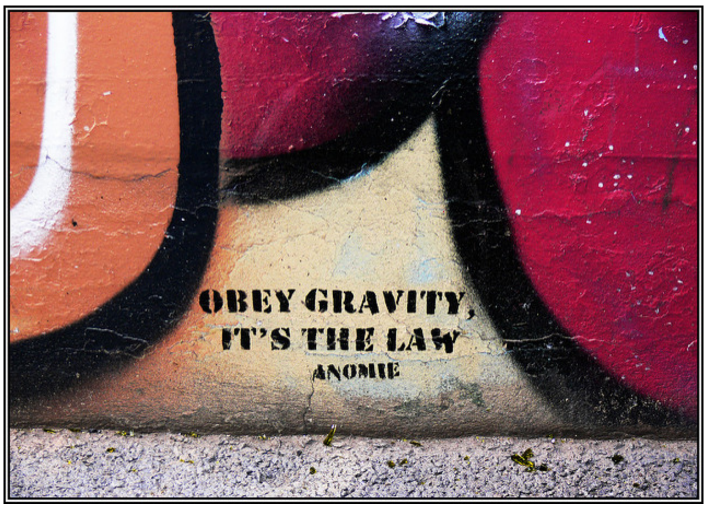


Figure 1.3.1: Defining a law. Does this graffiti mean "law" the way science defines "law"? (CC BY-SA-NC-ND; Koppenbadger).

There is an additional phrase in our definition of science: "the natural universe." Science is concerned *only* with the natural universe. What is the natural universe? It's anything that occurs around us, well, naturally. Stars, planets, the appearance of life on earth; as well as how animals, plants, and other matter function are all part of the natural universe. Science is concerned with that—and *only* that.

Of course, there are other things that concern us. For example, is the English language part of science? Most of us can easily answer no; English is not science. English is certainly worth knowing (at least for people in predominantly English-speaking countries), but why isn't it science? English, or any human language, is not science because ultimately it is *contrived*; it is made up. Think of it: the word spelled b-l-u-e represents a certain color, and we all agree what color that is. But what if we used the word h-a-r-d to describe that color? (Figure 1.3.2) That would be fine—as long as everyone agreed. Anyone who has learned a second language must initially wonder why a certain word is used to describe a certain concept; ultimately, the speakers of that language agreed that a particular word would represent a particular concept. It was contrived.

That doesn't mean language isn't worth knowing. It is very important in society. But it's not *science*. Science deals only with what occurs naturally.

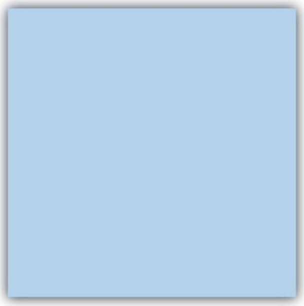


Figure 1.3.2: English Is Not Science. How would you describe this color? Blue or hard? Either way, you're not doing science.

Example 1.3.1: Identifying Science

Which of the following fields would be considered science?

a. geology, the study of the earth
b. ethics, the study of morality
c. political science, the study of governance
d. biology, the study of living organisms

Solution

a. Because the earth is a natural object, the study of it is indeed considered part of science.
b. Ethics is a branch of philosophy that deals with right and wrong. Although these are useful concepts, they are not science.
c. There are many forms of government, but all are created by humans. Despite the fact that the word *science* appears in its name, political science is not true science.
d. Living organisms are part of the natural universe, so the study of them is part of science.

Exercise 1.3.1

Which is part of science, and which is not?

a. dynamics, the study of systems that change over time
b. aesthetics, the concept of beauty

Answer A

science

Answer B

not science

The field of science has gotten so big that it is common to separate it into more specific fields. First, there is mathematics, the language of science. All scientific fields use mathematics to express themselves—some more than others. Physics and astronomy are scientific fields concerned with the fundamental interactions between matter and energy. Chemistry, as defined previously, is the study of the interactions of matter with other matter and with energy. Biology is the study of living organisms, while geology is the study of the earth. Other sciences can be named as well. Understand that these fields are not always completely separate; the boundaries between scientific fields are not always readily apparent. A scientist may be labeled a biochemist if he or she studies the chemistry of biological organisms.

Finally, understand that science can be either qualitative or quantitative. Qualitative implies a description of the quality of an object. For example, physical properties are generally qualitative descriptions: sulfur is yellow, your math book is heavy, or that statue is pretty. A quantitative description represents the specific amount of something; it means knowing how much of something is present, usually by counting or measuring it. Some quantitative descriptions include: 25 students in a class, 650 pages in a book, or a velocity of 66 miles per hour. Quantitative expressions are very important in science; they are also very important in chemistry.

Example 1.3.2: qualitative vs. quantitative Descriptions

Identify each statement as either a qualitative description or a quantitative description.

a. Gold metal is yellow.
b. A ream of paper has 500 sheets in it.
c. The weather outside is snowy.
d. The temperature outside is 24 degrees Fahrenheit.

Solution

a. Because we are describing a physical property of gold, this statement is qualitative.
b. This statement mentions a specific amount, so it is quantitative.
c. The word *snowy* is a description of how the day is; therefore, it is a qualitative statement.
d. In this case, the weather is described with a specific quantity—the temperature. Therefore, it is quantitative.

Exercise 1.3.2

Are these qualitative or quantitative statements?

a. Roses are red, and violets are blue.
b. Four score and seven years ago….

Answer A

qualitative

Answer B

quantitative

Food and Drink Application: Carbonated Beverages

Some of the simple chemical principles discussed in this chapter can be illustrated with carbonated beverages: sodas, beer, and sparkling wines. Each product is produced in a different way, but they all have one thing in common: they are solutions of carbon dioxide dissolved in water.

Carbon dioxide is a compound composed of carbon and oxygen. Under normal conditions, it is a gas. If you cool it down enough, it becomes a solid known as dry ice. Carbon dioxide is an important compound in the cycle of life on earth.

Even though it is a gas, carbon dioxide can dissolve in water, just like sugar or salt can dissolve in water. When that occurs, we have a homogeneous mixture, or a solution, of carbon dioxide in water. However, very little carbon dioxide can dissolve in water. If the atmosphere were pure carbon dioxide, the solution would be only about 0.07% carbon dioxide. In reality, the air is only about 0.03% carbon dioxide, so the amount of carbon dioxide in water is reduced proportionally.

However, when soda and beer are made, manufacturers do two important things: they use pure carbon dioxide gas, and they use it at very high pressures. With higher pressures, more carbon dioxide can dissolve in the water. When the soda or beer container is sealed, the high pressure of carbon dioxide gas remains inside the package. (Of course, there are more ingredients in soda and beer besides carbon dioxide and water.)

When you open a container of soda or beer, you hear a distinctive *hiss* as the excess carbon dioxide gas escapes. But something else happens as well. The carbon dioxide in the solution comes out of solution as a bunch of tiny bubbles. These bubbles impart a pleasing sensation in the mouth, so much so that the soda industry sold over *225 billion* servings of soda in the United States alone in 2009.

Some sparkling wines are made in the same way—by forcing carbon dioxide into regular wine. Some sparkling wines (including champagne) are made by sealing a bottle of wine with some yeast in it. The yeast *ferments*, a process by which the yeast converts sugars into energy and excess carbon dioxide. The carbon dioxide produced by the yeast dissolves in the wine. Then, when the champagne bottle is opened, the increased pressure of carbon dioxide is released, and the drink bubbles just like an expensive glass of soda.

Figure 1.3.3: Carbonated Beverages © Thinkstock. Soda, beer, and sparkling wine take advantage of the properties of a solution of carbon dioxide in water.

Soda (left), beer (middle) and sparkling wine (right) images next to one another.

Key Takeaways

- Science is a process of knowing about the natural universe through observation and experiment.
- Scientists go through a rigorous process to determine new knowledge about the universe; this process is generally referred to as the scientific method.
- Science is broken down into various fields, of which chemistry is one.
- Science, including chemistry, is both qualitative and quantitative.

CHAPTER OVERVIEW

2: PSS- Scientific Measurements

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

2.1: PSS.1- Uncertainty in Measurements

Learning Objective

- Learn the units that go with various quantities
- Express units using their abbreviations
- Make new units by combining numerical prefixes with units

A number indicates "how much," but the unit indicates "of what." The "of what" is important when communicating a quantity. For example, if you were to ask a friend how close you are to Lake Erie and your friend says "six," then your friend isn't giving you complete information. Six *what*? Six miles? Six inches? Six city blocks? The actual distance to the lake depends on what units you use.

Chemistry, like most sciences, uses the International System of Units, or SI for short. (The letters *SI* stand for the French "le Système International d'unités.") SI specifies certain units for various types of quantities, based on seven fundamental units. We will use most of the fundamental units in chemistry. Initially, we will deal with three fundamental units. The **meter** (m) is the SI unit of length. It is a little longer than a yard (Figure 2.1.1). The SI unit of mass is the **kilogram** (kg), which is about 2.2 pounds (lb). The SI unit of time is the **second** (s).

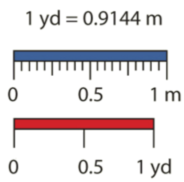


Figure 2.1.1: The Meter. The SI standard unit of length, the meter, is a little longer than a yard.

To express a quantity, you need to combine a number with a unit. If you have a length that is 2.4 m, then you express that length as simply 2.4 m. A time of 15,000 s can be expressed as 1.5×10^4 s in scientific notation.

Sometimes, a given unit is not an appropriate size to easily express a quantity. For example, the width of a human hair is very small, and it doesn't make much sense to express it in meters. SI also defines a series of *numerical prefixes*, referring to multiples or fractions of a fundamental unit, to make a unit more conveniently sized for a specific quantity. Table 2.1.1 lists the prefixes, their abbreviations, and their multiplicative factors. Some of the prefixes, such as kilo-, mega-, and giga-, represent more than one of the fundamental unit, while other prefixes, such as centi-, milli-, and micro-, represent fractions of the original unit. Note, too, that once again we are using powers of 10. Each prefix is a multiple of or fraction of a power of 10.

Table 2.1.1: Multiplicative Prefixes for SI Units

Prefix	Abbreviation	Multiplicative Amount	Power of Ten
giga-	G	1,000,000,000 ×	10^9 ×
mega-	M	1,000,000 ×	10^6 ×
kilo-	k	1,000 ×	10^3 ×
deci-	d	1/10 ×	×
centi-	c	1/100 ×	×
milli-	m	1/1,000 ×	×
micro-	µ*	1/1,000,000 ×	×
nano-	n	1/1,000,000,000 ×	×
pico-	p	1/1,000,000,000,000 ×	x
* The letter *μ* is the Greek letter lowercase equivalent to an m and is called "mu" (pronounced "myoo").			

To use the fractions to generate new units, simply combine the prefix with the unit itself; the abbreviation for the new unit is the combination of the abbreviation for the prefix and the abbreviation of the unit. For example, the kilometer (km) is 1,000 × meter, or 1,000 m. Thus, 5 kilometers (5 km) is equal to 5,000 m. Similarly, a millisecond (ms) is 1/1,000 × second, or one-thousandth of a second. Thus, 25 ms is 25 thousandths of a second. You will need to become proficient in combining prefixes and units. (You may recognize that one of our fundamental units, the kilogram, automatically has a prefix-unit combination. The word *kilogram* means 1,000 g.)

In addition to the fundamental units, SI also allows for derived units based on a fundamental unit or units. There are many derived units used in science. For example, the derived unit for area comes from the idea that area is defined as width times height. Because both width and height are lengths, they both have the fundamental unit of meter, so the unit of area is meter × meter, or meter^2 (m^2). This is sometimes spoken as "square meters." A unit with a prefix can also be used to derive a unit for area, so we can also have cm^2, mm^2, or km^2 as acceptable units for area.

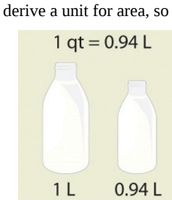


Figure 2.1.2: The Liter. The SI unit of volume, the liter, is slightly larger than 1 quart.

Volume is defined as length times width times height, so it has units of meter × meter × meter, or meter^3 (m^3)—sometimes spoken as "cubic meters." The cubic meter is a rather large unit, however, so another unit is defined that is somewhat more manageable: the liter (L). A liter is 1/1,000th of a cubic meter and is a little more than 1 quart in volume (Figure 2.1.2). Prefixes can also be used with the liter unit, so we can speak of milliliters (1/1,000th of a liter; mL) and kiloliters (1,000 L; kL).

Another definition of a liter is one-tenth of a meter cubed. Because one-tenth of a meter is 10 cm, then a liter is equal to 1,000 cm^3 (Figure 2.1.3). Because 1 L equals 1,000 mL, we conclude that 1 mL equals 1 cm^3; thus, these units are interchangeable.

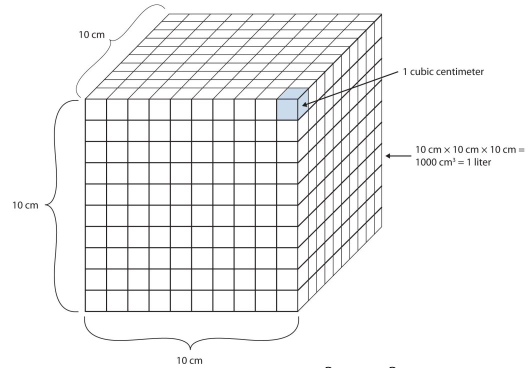


Figure 2.1.3: The size of one liter equals 1,000 cm^3, so 1 cm^3 is the same as 1 mL.

Units are not only multiplied together—they can also be divided. For example, if you are traveling at one meter for every second of time elapsed, your velocity is 1 meter per second, or 1 m/s. The word *per* implies division, so velocity is determined by dividing a distance quantity by a time quantity. Other units for velocity include kilometers per hour (km/h) or even micrometers per nanosecond (μm/ns). Later, we will see other derived units that can be expressed as fractions.

Example 2.1.1

a. A human hair has a diameter of about 6.0×10^{-5} m. Suggest an appropriate unit for this measurement and write the diameter of a human hair in terms of that unit.
b. What is the velocity of a car if it goes 25 m in 5.0 s?

Solution

a. The scientific notation 10^{-5} is close to 10^{-6}, which defines the micro- prefix. Let us use micrometers as the unit for hair diameter. The number 6.0×10^{-5} can be written as 60×10^{-6}, and a micrometer is 10^{-6} m, so the diameter of a human hair is about 60 μm.
b. If velocity is defined as a distance quantity divided by a time quantity, then velocity is 25 meters/5.0 seconds. Dividing the numbers gives us 25/5.0 = 5.0, and dividing the units gives us meters/second, or m/s. The velocity is 5.0 m/s.

Exercise 2.1.1

a. Express the volume of an Olympic-sized swimming pool, 2,500,000 L, in more appropriate units.
b. A common garden snail moves about 6.1 m in 30 min. What is its velocity in meters per minute (m/min)?

Answer a

- 2.5 ML

Answer b

- 0.203 m/min

Key Takeaways

- Numbers tell "how much," and units tell "of what."
- Chemistry uses a set of fundamental units and derived units from SI units.
- Chemistry uses a set of prefixes that represent multiples or fractions of units.
- Units can be multiplied and divided to generate new units for quantities.

2.2: PSS.2- Significant Digits

Learning Objectives

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

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

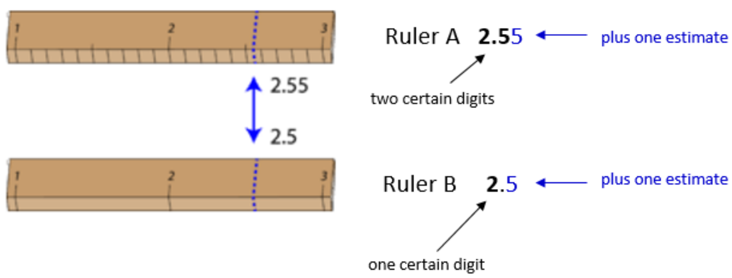


Figure $2.2.1$: Measurement with two different rulers.
Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

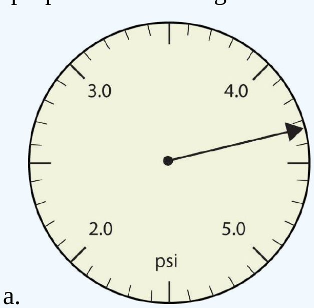
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest $0.1\ \text{g}$, other highly sensitive balances are capable of measuring to the nearest $0.001\ \text{g}$ or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure $2.2.1$ shows two rulers making the same measurement of an object (indicated by the blue arrow).

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

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

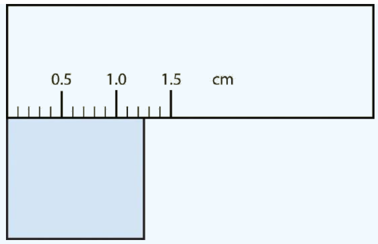
Example $2.2.1$: Reporting Measurements to the Proper Number of Significant Figures

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



a.

b.



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

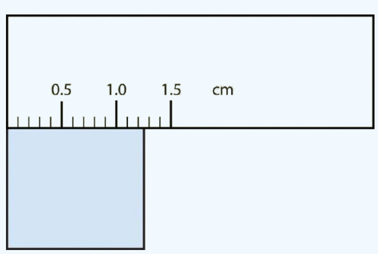
Solutions

Solutions to Example 2.3.1

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

? Exercise 2.2.1

What would be the reported width of this rectangle?



Answer

1.25 cm

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

Table 2.2.1: Significant Figure Rules

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

Exact Numbers

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

✔ Example 2.2.2

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

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

Solution

Solution to Example 2.3.2

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures

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

? Exercise 2.2.2

Give the number of significant figures in each.

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

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

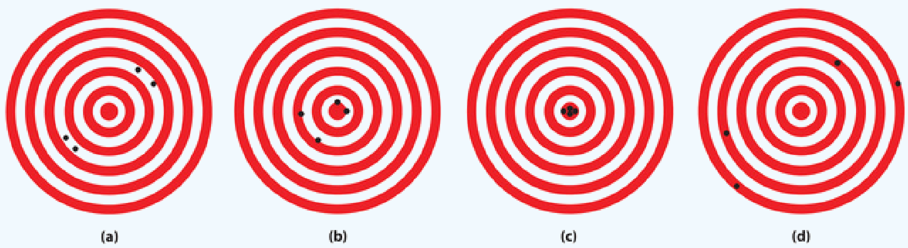
Answer e

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

Accuracy and Precision

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

Video 2.2.1: Difference between precision and accuracy.

> **Example 2.2.3**
>
> The following archery targets show marks that represent the results of four sets of measurements.
>
> 
>
>
> Which target shows
>
> a. a precise, but inaccurate set of measurements?
> b. a set of measurements that is both precise and accurate?
> c. a set of measurements that is neither precise nor accurate?
>
> **Solution**
>
> a. Set **(a)** is precise, but inaccurate.
> b. Set **(c)** is both precise and accurate.
> c. Set **(d)** is neither precise nor accurate.

Summary

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

2.3: PSS.3- Rounding Off Nonsignificant Digits

Learning Objectives

- Give an example of a measurement whose number of significant digits is clearly too great, and explain why.
- State the purpose of rounding off, and describe the information that must be known to do it properly.
- Round off a number to a specified number of significant digits.
- Explain how to round off a number whose second-most-significant digit is 9.
- Carry out a simple calculation that involves two or more observed quantities, and express the result in the appropriate number of significant figures.

The numerical values we deal with in science (and in many other aspects of life) represent measurements whose values are never known exactly. Our pocket-calculators or computers don't know this; they treat the numbers we punch into them as "pure" mathematical entities, with the result that the operations of arithmetic frequently yield answers that are physically ridiculous even though mathematically correct. The purpose of this unit is to help you understand why this happens, and to show you what to do about it.

Digits: Significant and otherwise

Consider the two statements shown below:

- **"The population of our city is 157,872."**
- **"The number of registered voters as of Jan 1 was 27,833.**

Which of these would you be justified in dismissing immediately? Certainly not the second one, because it probably comes from a database which contains one record for each voter, so the number is found simply by counting the number of records. The first statement cannot possibly be correct. Even if a city's population could be defined in a precise way (Permanent residents? Warm bodies?), how can we account for the minute-by minute changes that occur as people are born and die, or move in and move away?

What is the difference between the two population numbers stated above? The first one expresses a quantity that cannot be known exactly — that is, it carries with it a degree of uncertainty. It is quite possible that the last census yielded precisely 157,872 records, and that this might be the "population of the city" for legal purposes, but it is surely not the "true" population. To better reflect this fact, one might list the population (in an atlas, for example) as **157,900** or even **158,000**. These two quantities have been rounded off to four and three significant figures, respectively, and the have the following meanings:

- **157900** (the significant digits are underlined here) implies that the population is believed to be within the range of about 157850 to about 157950. In other words, the population is 157900±50. The "plus-or-minus 50" appended to this number means that we consider the absolute uncertainty of the population measurement to be 50 – (–50) = 100. We can also say that the relative uncertainty is 100/157900, which we can also express as 1 part in 1579, or 1/1579 = 0.000633, or about 0.06 percent.
- The value **158000** implies that the population is likely between about 157500 and 158500, or 158000±500. The absolute uncertainty of 1000 translates into a relative uncertainty of 1000/158000 or 1 part in 158, or about 0.6 percent.

Which of these two values we would report as "the population" will depend on the degree of confidence we have in the original census figure; if the census was completed last week, we might round to four significant digits, but if it was a year or so ago, rounding to three places might be a more prudent choice. In a case such as this, there is no really objective way of choosing between the two alternatives.

This illustrates an important point: the concept of *significant digits* has less to do with mathematics than with our confidence in a measurement. This confidence can often be expressed numerically (for example, the height of a liquid in a measuring tube can be read to ±0.05 cm), but when it cannot, as in our population example, we must depend on our personal experience and judgment.

So, what is a significant digit? According to the usual definition, it is all the numerals in a measured quantity (counting from the left) whose values are considered as known exactly, plus one more whose value could be one more or one less:

- In "157900" (four significant digits), the left most three digits are known exactly, but the fourth digit, "9" could well be "8" if the "true value" is within the implied range of 157850 to 157950.
- In "*158*000" (three significant digits), the left most two digits are known exactly, while the third digit could be either "7" or "8" if the true value is within the implied range of 157500 to 158500.

Although rounding off always leads to the loss of numeric information, what we are getting rid of can be considered to be "numeric noise" that does not contribute to the quality of the measurement. The purpose in rounding off is to avoid expressing a value to a greater degree of precision than is consistent with the uncertainty in the measurement.

Implied Uncertainty

If you know that a balance is accurate to within 0.1 mg, say, then the uncertainty in any measurement of mass carried out on this balance will be ±0.1 mg. Suppose, however, that you are simply told that an object has a length of 0.42 cm, with no indication of its precision. In this case, all you have to go on is the number of digits contained in the data. Thus the quantity "0.42 cm" is specified to 0.01 unit in 0 42, or one part in 42 . The implied relative uncertainty in this figure is 1/42, or about 2%. The precision of any numeric answer calculated from this value is therefore limited to about the same amount.

Rounding Error

It is important to understand that the number of significant digits in a value provides only a rough indication of its precision, and that information is lost when rounding off occurs. Suppose, for example, that we measure the weight of an object as 3.28 g on a balance believed to be accurate to within ±0.05 gram. The resulting value of 3.28±.05 gram tells us that the true weight of the object could be anywhere between 3.23 g and 3.33 g. The absolute uncertainty here is 0.1 g (±0.05 g), and the relative uncertainty is 1 part in 32.8, or about 3 percent.

How many significant digits should there be in the reported measurement? Since only the left most "3" in "3.28" is certain, you would probably elect to round the value to 3.3 g. So far, so good. But what is someone else supposed to make of this figure when they see it in your report? The value "3.3 g" suggests an *implied uncertainty* of 3.3±0.05 g, meaning that the true value is likely between 3.25 g and 3.35 g. This range is 0.02 g below that associated with the original measurement, and so rounding off has introduced a bias of this amount into the result. Since this is less than half of the ±0.05 g uncertainty in the weighing, it is not a very serious matter in itself. However, if several values that were rounded in this way are combined in a calculation, the rounding-off errors could become significant.

Rules for Rounding

The standard rules for rounding off are well known. Before we set them out, let us agree on what to call the various components of a numeric value.

- The *most significant digit* is the left most digit (not counting any leading zeros which function only as placeholders and are never significant digits.)
- If you are rounding off to ***n*** significant digits, then the *least significant digit* is the ***n***th digit from the most significant digit. The least significant digit can be a zero.
- The first non-significant digit is the ***n***+1th digit.

Rounding-off rules

- If the first non-significant digit is less than 5, then the least significant digit remains unchanged.
- If the first non-significant digit is greater than 5, the least significant digit is incremented by 1.
- If the first non-significant digit is 5, the least significant digit can either be incremented or left unchanged (*see below!*)
- All non-significant digits are removed.

Fantasies about fives

Students are sometimes told to increment the least significant digit by 1 if it is odd, and to leave it unchanged if it is even. One wonders if this reflects some idea that even numbers are somehow “better” than odd ones! (The ancient superstition is just the opposite, that only the odd numbers are "lucky".)

In fact, you could do it equally the other way around, incrementing only the even numbers. If you are only rounding a single number, it doesn't really matter what you do. However, when you are rounding a series of numbers that will be used in a calculation, if you treated each first nonsignificant 5 in the same way, you would be over- or understating the value of the rounded number, thus accumulating round-off error. Since there are equal numbers of even and odd digits, incrementing only the one kind will keep this kind of error from building up. You could do just as well, of course, by flipping a coin!

Table 2.3.1: Examples of rounding-off

number to round	number of significant digits	result	comment
34.216	3	34.2	First non-significant digit (1) is less than 5, so number is simply truncated.
2.252	2	2.2 or 2.3	First non-significant digit is 5, so least sig. digit can either remain unchanged or be incremented.
39.99	3	40.0	Crossing "decimal boundary", so all numbers change.
85,381	3	85,400	The two zeros are just placeholders
0.04597	3	0.0460	The two leading zeros are not significant digits.

Rounding up the Nines

Suppose that an object is found to have a weight of 3.98 ± 0.05 g. This would place its true weight somewhere in the range of **3.93** g to **4.03** g. In judging how to round this number, you count the number of digits in “3.98” that are known exactly, and you find none! Since the “4” is the left most digit whose value is uncertain, this would imply that the result should be rounded to one significant figure and reported simply as 4 g. An alternative would be to bend the rule and round off to two significant digits, yielding 4.0 g. How can you decide what to do? In a case such as this, you should look at the implied uncertainties in the two values, and compare them with the uncertainty associated with the original measurement.

Table 2.3.2

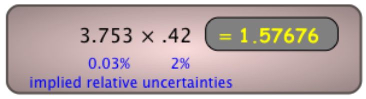
rounded value	implied max	implied min	absolute uncertainty	relative uncertainty
3.98 g	3.985 g	3.975 g	±.005 g or 0.01 g	1 in 400, or 0.25%
4 g	4.5 g	3.5 g	±.5 g or 1 g	1 in 4, 25%
4.0 g	**4.05 g**	**3.95 g**	**±.05 g or 0.1 g**	**1 in 40, 2.5%**

Clearly, rounding off to two digits is the only reasonable course in this example. Observed values should be rounded off to the number of digits that most accurately conveys the uncertainty in the measurement.

- Usually, this means rounding off to the number of significant digits in in the quantity; that is, the number of digits (counting from the left) that are known exactly, plus one more.
- When this cannot be applied (as in the example above when addition of subtraction of the absolute uncertainty bridges a power of ten), then we round in such a way that the relative implied uncertainty in the result is as close as possible to that of the observed value.

Rounding the Results of Calculations

When carrying out calculations that involve multiple steps, you should avoid doing any rounding until you obtain the final result. Suppose you use your calculator to work out the area of a rectangle:



rounded value	relative implied uncertainty
1.58	1 part in 158, or 0.6%
1.6	1 part in 16, or 6 %

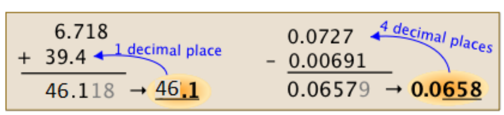
> **Note**
>
> Your calculator is of course correct as far as the pure numbers go, but you would be wrong to write down "1.57676 cm^2" as the answer. Two possible options for rounding off the calculator answer are shown at the right.

It is clear that neither option is entirely satisfactory; rounding to 3 significant digits overstates the precision of the answer, whereas following the rule and rounding to the two digits in ".42" has the effect of throwing away some precision. In this case, it could be argued that rounding to three digits is justified because the implied relative uncertainty in the answer, 0.6%, is more consistent with those of the two factors.

> *The "rules" for rounding off are generally useful, convenient guidelines, but they do not always yield the most desirable result. When in doubt, it is better to rely on relative implied uncertainties.*

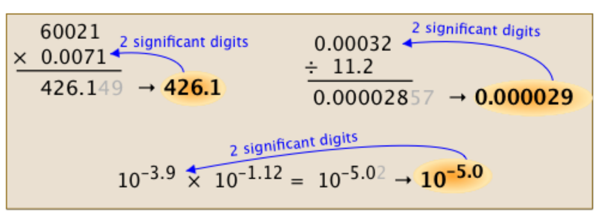
Addition and Subtraction

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 that the least precise number used to get the answer. When adding or subtracting, we go by the number of *decimal places* (i.e., the number of digits on the right side of the decimal point) rather than by the number of significant digits. Identify the quantity having the smallest number of decimal places, and use this number to set the number of decimal places in the answer.



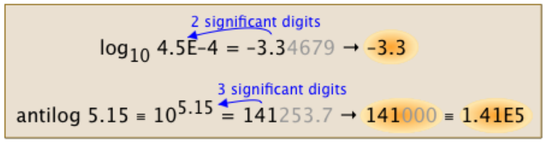
Multiplication and Division

The result must contain the same number of significant figures as in the value having the least number of significant figures.



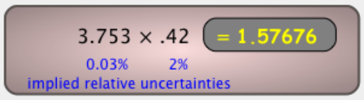
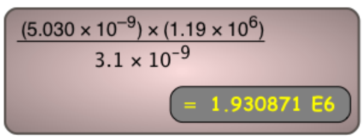
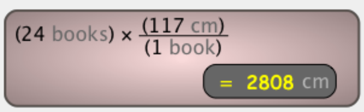
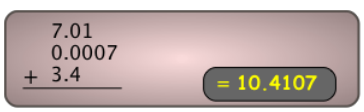
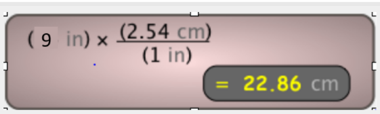
Logarithms and antilogarithms

If a number is expressed in the form $a \times 10^b$ ("scientific notation") with the additional restriction that the coefficient a is no less than 1 and less than 10, the number is in its *normalized* form. Express the base-10 logarithm of a value using the same number of significant figures as is present in the *normalized form* of that value. Similarly, for antilogarithms (numbers expressed as powers of 10), use the same number of significant figures as are in that power.

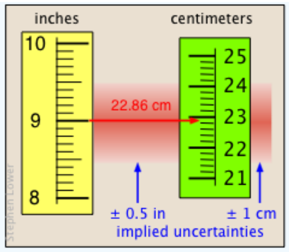


Examples 2.3.1

The following examples will illustrate the most common problems you are likely to encounter in rounding off the results of calculations. They deserve your careful study!

calculator result	rounded	remarks
	1.6	Rounding to two significant figures yields an implied uncertainty of 1/16 or 6%, three times greater than that in the least-precisely known factor. This is a good illustration of how rounding can lead to the loss of information.
	1.9E6	The "3.1" factor is specified to 1 part in 31, or 3%. In the answer 1.9, the value is expressed to 1 part in 19, or 5%. These precisions are comparable, so the rounding-off rule has given us a reasonable result.
A certain book has a thickness of 117 mm; find the height of a stack of 24 identical books: 	2810 mm	The “24” and the “1” are exact, so the only uncertain value is the thickness of each book, given to 3 significant digits. The trailing zero in the answer is only a placeholder.
	10.4	In addition or subtraction, look for the term having the smallest number of decimal places, and round off the answer to the same number of places.
		
	23 cm	see below

The last of the examples shown above represents the very common operation of converting one unit into another. There is a certain amount of ambiguity here; if we take "9 in" to mean a distance in the range 8.5 to 9.5 inches, then the implied uncertainty is ±0.5 in, which is 1 part in 18, or about ± 6%. The relative uncertainty in the answer must be the same, since all the values are multiplied by the same factor, 2.54 cm/in. In this case we are justified in writing the answer to two significant digits, yielding an uncertainty of about ±1 cm; if we had used the answer "20 cm" (one significant digit), its implied uncertainty would be ±5 cm, or ±25%.



When the appropriate number of significant digits is in question, calculating the relative uncertainty can help you decide.

2.4: PSS.4- Adding and Subtracting Measurements

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2.5: PSS.5- Multiplying and Dividing Measurements

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2.6: PSS.6- Exponential Numbers

Learning Objectives

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

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

$$N \times 10^n$$

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

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

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

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example $2.6.1$.

Example $2.6.1$: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

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

Solution

Solutions to Example 2.2.1

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

	Explanation	Answer
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^{1}
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^{3}
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example $2.6.2$ illustrates how to do this.

✔ Example 2.6.2: Expressing Sums and Differences in Scientific Notation

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

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

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^{2}) + (47.3 \times 10^{2}) =$ $(1.36 + 47.3) \times 10^{2} = 48.66 \times 10^{2}$ or $(0.136 \times 10^{3}) + (4.73 \times 10^{3}) =$ $(0.136 + 4.73) \times 10^{3}) = 4.87 \times 10^{3}$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^{2} to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.8
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.0

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example $2.6.3$.

Example $2.6.3$: Expressing Products and Quotients in Scientific Notation

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

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

Solution

Solution to Example 2.2.3

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

2.7: PSS.7- Scientific Notation

Learning Objectives

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

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

$$N \times 10^n$$

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

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

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Example 2.7.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

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

Solution

Solutions to Example 2.2.1

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

	Explanation	Answer
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^{1}
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^{3}
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

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Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example $2.7.2$ illustrates how to do this.

✔ Example 2.7.2: Expressing Sums and Differences in Scientific Notation

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

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

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) =$ $(1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) =$ $(0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.8
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.0

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example $2.7.3$.

Example 2.7.3: Expressing Products and Quotients in Scientific Notation

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

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

Solution

Solution to Example 2.2.3

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

2.8: PSS.8- Unit Equations and Unit Factors

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2.9: PSS.9- Unit Analysis Problem Solving

Learning Objective

- Convert from one unit to another unit of the same type.

In Section 2.2, we showed some examples of how to replace initial units with other units of the same type to get a numerical value that is easier to comprehend. In this section, we will formalize the process.

Consider a simple example: how many feet are there in 4 yards? Most people will almost automatically answer that there are 12 feet in 4 yards. How did you make this determination? Well, if there are 3 feet in 1 yard and there are 4 yards, then there are 4 × 3 = 12 feet in 4 yards.

This is correct, of course, but it is informal. Let us formalize it in a way that can be applied more generally. We know that 1 yard (yd) equals 3 feet (ft):

$$1\,yd = 3\,ft$$

In math, this expression is called an *equality*. The rules of algebra say that you can change (i.e., multiply or divide or add or subtract) the equality (as long as you do not divide by zero) and the new expression will still be an equality. For example, if we divide both sides by 2, we get:

$$\frac{1}{2}\,yd = \frac{3}{2}\,ft$$

We see that one-half of a yard equals 3/2, or one and a half, feet—something we also know to be true—so the above equation is still an equality. Going back to the original equality, suppose we divide both sides of the equation by 1 yard (number **and** unit):

$$\frac{1\,yd}{1\,yd} = \frac{3\,ft}{1\,yd}$$

The expression is still an equality, by the rules of algebra. The left fraction equals 1. It has the same quantity in the numerator and the denominator, so it must equal 1. The quantities in the numerator and denominator cancel, both the number *and* the unit:

$$\frac{1\,yd}{1\,yd} = \frac{3\,ft}{1\,yd}$$

When everything cancels in a fraction, the fraction reduces to 1:

$$1 = \frac{3\,ft}{1\,yd}$$

Conversion Factors

We have an expression that equals 1.

$$\frac{3\,ft}{1\,yd} = 1$$

This is a strange way to write 1, but it makes sense: 3 ft equal 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units.

The expression

$$\frac{3\,ft}{1\,yd}$$

is called a conversion factor and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

To see how this happens, let us start with the original quantity: 4 yd.

Now let us multiply this quantity by 1. When you multiply anything by 1, you do not change the value of the quantity. Rather than multiplying by just 1, let us write 1 as:

$$\frac{3\,ft}{1\,yd}$$

$$4\,yd \times \frac{3\,ft}{1\,yd}$$

The 4 yd term can be thought of as 4yd/1; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. In this case, what cancels is the unit *yard*:

$$4\,yd \times \frac{3\,ft}{1\,yd}$$

That is all that we can cancel. Now, multiply and divide all the numbers to get the final answer:

$$\frac{4 \times 3\,ft}{1} = \frac{12\,ft}{1} = 12\,ft$$

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is applicable to a variety of problems.

How many millimeters are in 14.66 m? To answer this, we need to construct a conversion factor between millimeters and meters and apply it correctly to the original quantity. We start with the definition of a millimeter, which is:

$$1\,mm = \frac{1}{1000\,m}$$

The 1/1000 is what the prefix *milli*- means. Most people are more comfortable working without fractions, so we will rewrite this equation by bringing the 1,000 into the numerator of the other side of the equation:

$$1000\,mm = 1\,m$$

Now we construct a conversion factor by dividing one quantity into both sides. But now a question arises: which quantity do we divide by? It turns out that we have two choices, and the two choices will give us different conversion factors, both of which equal 1:

$$\frac{1000\,mm}{1000\,mm} = \frac{1\,m}{1000\,mm}$$

or

$$\frac{1000\,mm}{1\,m} = \frac{1\,m}{1\,m}$$

$$1 = \frac{1\,m}{1000\,mm}$$

or

$$\frac{1000\,mm}{1\,m} = 1$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is meters, which we want to convert to millimeters. Because the original unit is assumed to be in the numerator, to get rid of it, we want the meter unit in the *denominator*; then they will cancel. Therefore, we will use the second conversion factor. Canceling units and performing the mathematics, we get:

$$14.66m \times \frac{1000\,mm}{1\,m} = 14660\,mm$$

Note how m cancels, leaving mm, which is the unit of interest.

The ability to construct and apply proper conversion factors is a very powerful mathematical technique in chemistry. You need to master this technique if you are going to be successful in this and future courses.

Example 2.9.1

a. Convert 35.9 kL to liters.
b. Convert 555 nm to meters.

Solution

a. We will use the fact that 1 kL = 1,000 L. Of the two conversion factors that can be defined, the one that will work is 1000L/1kL. Applying this conversion factor, we get:

$$35.9\,kL \times \frac{1000\,L}{1\,kL} = 35,900\,L$$

b. We will use the fact that 1 nm = 1/1,000,000,000 m, which we will rewrite as 1,000,000,000 nm = 1 m, or 10^9 nm = 1 m. Of the two possible conversion factors, the appropriate one has the nm unit in the denominator:

$$\frac{1\,m}{10^9\,nm}$$

Applying this conversion factor, we get:

$$555\,nm \times \frac{1m}{10^9 nm} = 0.000000555\,m = 5.55 \times 10^{-7}\,m$$

In the final step, we expressed the answer in scientific notation.

Exercise 2.9.1

a. Convert 67.08 μL to liters.
b. Convert 56.8 m to kilometers.

Answer a

6.708×10^{-5} L

Answer b

5.68×10^{-2} km

What if we have a derived unit that is the product of more than one unit, such as m^2? Suppose we want to convert square meters to square centimeters? The key is to remember that m^2 means m × m, which means we have *two* meter units in our derived unit. That means we have to include *two* conversion factors, one for each unit. For example, to convert 17.6 m^2 to square centimeters, we perform the conversion as follows:

$$\begin{aligned} 17.6m^2 &= 17.6(m \times m) \times \frac{100cm}{1m} \times \frac{100cm}{1m} \\ &= 176000\,cm \times cm \\ &= 1.76 \times 10^5\,cm^2 \end{aligned} \tag{2.9.1}$$

Example 2.9.2

How many cubic centimeters are in 0.883 m^3?

Solution

With an exponent of 3, we have three length units, so by extension we need to use three conversion factors between meters and centimeters. Thus, we have:

$$0.883m^3 \times \frac{100\,cm}{1\,m} \times \frac{100\,cm}{1\,m} \times \frac{100\,cm}{1\,m} = 883000\,cm^3 = 8.83 \times 10^5\,cm^3$$

You should demonstrate to yourself that the three meter units do indeed cancel.

Exercise 2.9.2

How many cubic millimeters are present in 0.0923 m^3?

Answer

9.23×10^7 mm^3

Suppose the unit you want to convert is in the denominator of a derived unit—what then? Then, in the conversion factor, the unit you want to remove must be in the *numerator*. This will cancel with the original unit in the denominator and introduce a new unit in the denominator. The following example illustrates this situation.

Example 2.9.3

Convert 88.4 m/min to meters/second.

Solution

We want to change the unit in the denominator from minutes to seconds. Because there are 60 seconds in 1 minute (60 s = 1 min), we construct a conversion factor so that the unit we want to remove, minutes, is in the numerator: 1min/60s. Apply and perform the math:

$$\frac{88.4m}{min} \times \frac{1\,min}{60\,s} = 1.47\frac{m}{s}$$

Notice how the 88.4 automatically goes in the numerator. That's because any number can be thought of as being in the numerator of a fraction divided by 1.

Exercise 2.9.3

Convert 0.203 m/min to meters/second.

Answer

0.00338 m/s

or

3.38×10^{-3} m/s

Sometimes there will be a need to convert from one unit with one numerical prefix to another unit with a different numerical prefix. How do we handle those conversions? Well, you could memorize the conversion factors that interrelate all numerical prefixes. Or you can go the easier route: first convert the quantity to the base unit—the unit with no numerical prefix—using the definition of the original prefix. Then, convert the quantity in the base unit to the desired unit using the definition of the second prefix. You can do the conversion in two separate steps or as one long algebraic step. For example, to convert 2.77 kg to milligrams:

$$2.77\,kg \times \frac{1000\,g}{1\,kg} = 2770\,g$$

(convert to the base units of grams)

$$2770\,g \times \frac{1000\,mg}{1\,g} = 2770000\,mg = 2.77 \times 10^6\,mg$$

(convert to desired unit)

Alternatively, it can be done in a single multi-step process:

$$\begin{aligned} 2.77\ \cancel{kg} \times \frac{1000\ \cancel{g}}{1\ \cancel{kg}} \times \frac{1000\, mg}{1\ \cancel{g}} &= 2770000\, mg \\ &= 2.77 \times 10^6\, mg \end{aligned} \tag{2.9.2}$$

You get the same answer either way.

Example 2.9.4

How many nanoseconds are in 368.09 µs?

Solution

You can either do this as a one-step conversion from microseconds to nanoseconds or convert to the base unit first and then to the final desired unit. We will use the second method here, showing the two steps in a single line. Using the definitions of the prefixes *micro-* and *nano-*,

$$368.0\, \mu s \times \frac{1\ s}{1000000\, \mu s} \times \frac{1000000000}{1\ s} = 3.6809 \times 10^5\, ns$$

Exercise 2.9.4

How many milliliters are in 607.8 kL?

Answer

6.078×10^8 mL

When considering the significant figures of a final numerical answer in a conversion, there is one important case where a number does not impact the number of significant figures in a final answer: the so-called **exact number**. An exact number is a number from a defined relationship, not a measured one. For example, the prefix *kilo-* means 1,000-*exactly* 1,000, no more or no less. Thus, in constructing the conversion factor:

$$\frac{1000\, g}{1\, kg}$$

neither the 1,000 nor the 1 enter into our consideration of significant figures. The numbers in the numerator and denominator are defined exactly by what the prefix *kilo-* means. Another way of thinking about it is that these numbers can be thought of as having an infinite number of significant figures, such as:

$$\frac{1000.0000000000\ldots\ g}{1.0000000000\ldots\ kg}$$

The other numbers in the calculation will determine the number of significant figures in the final answer.

Example 2.9.5

A rectangular plot in a garden has the dimensions 36.7 cm by 128.8 cm. What is the area of the garden plot in square meters? Express your answer in the proper number of significant figures.

Solution

Area is defined as the product of the two dimensions, which we then have to convert to square meters, and express our final answer to the correct number of significant figures—which in this case will be three.

$$36.7\, cm \times 128.8\, cm \times \frac{1\, m}{100\, cm} \times \frac{1\, m}{100\, cm} = 0.472696\, m^2 = 0.473\, m^2$$

The 1 and 100 in the conversion factors do not affect the determination of significant figures because they are exact numbers, defined by the centi- prefix.

Exercise 2.9.5

What is the volume of a block in cubic meters with the dimensions 2.1 cm × 34.0 cm × 118 cm?

Answer

0.0084 m^3

Chemistry is Everywhere: The Gimli Glider

On July 23, 1983, an Air Canada Boeing 767 jet had to glide to an emergency landing at Gimli Industrial Park Airport in Gimli, Manitoba, because it unexpectedly ran out of fuel during flight. There was no loss of life in the course of the emergency landing, only some minor injuries associated in part with the evacuation of the craft after landing. For the remainder of its operational life (the plane was retired in 2008), the aircraft was nicknamed "the Gimli Glider."

The Gimli Glider is the Boeing 767 that ran out of fuel and glided to safety at Gimli Airport. The aircraft ran out of fuel because of confusion over the units used to express the amount of fuel. Source: Photo courtesy of Will F., (CC BY-SA 2.5; Aero Icarus).

The 767 took off from Montreal on its way to Ottawa, ultimately heading for Edmonton, Canada. About halfway through the flight, all the engines on the plane began to shut down because of a lack of fuel. When the final engine cut off, all electricity (which was generated by the engines) was lost; the plane became, essentially, a powerless glider. Captain Robert Pearson was an experienced glider pilot, although he had never flown a glider the size of a 767. First Officer Maurice Quintal quickly determined that the aircraft would not be able make it to Winnipeg, the next large airport. He suggested his old Royal Air Force base at Gimli Station, one of whose runways was still being used as a community airport. Between the efforts of the pilots and the flight crew, they managed to get the airplane safely on the ground (although with buckled landing gear) and all passengers off safely.

What happened? At the time, Canada was transitioning from the older English system to the metric system. The Boeing 767s were the first aircraft whose gauges were calibrated in the metric system of units (liters and kilograms) rather than the English system of units (gallons and pounds). Thus, when the fuel gauge read 22,300, the gauge meant kilograms, but the ground crew mistakenly fueled the plane with 22,300 *pounds* of fuel. This ended up being just less than half of the fuel needed to make the trip, causing the engines to quit about halfway to Ottawa. Quick thinking and extraordinary skill saved the lives of 61 passengers and 8 crew members—an incident that would not have occurred if people were watching their units.

Key Takeaways

- Units can be converted to other units using the proper conversion factors.
- Conversion factors are constructed from equalities that relate two different units.
- Conversions can be a single step or multi-step.
- Unit conversion is a powerful mathematical technique in chemistry that must be mastered.
- Exact numbers do not affect the determination of significant figures.

- **2.5: Converting Units** by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/beginning-chemistry.

2.10: PSS.10- The Percent Concept

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CHAPTER OVERVIEW

3: The Metric System

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

- **Homework**

Template:HideTOC

3.1: 2.1- Basic Units and Symbols

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

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

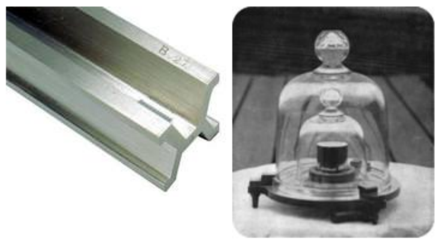

Figure 3.1.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 3.1.1.

Table 3.1.1: SI Base Units of Measurement

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

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

Unfortunately, the Metric System is Not Ubiquitous

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

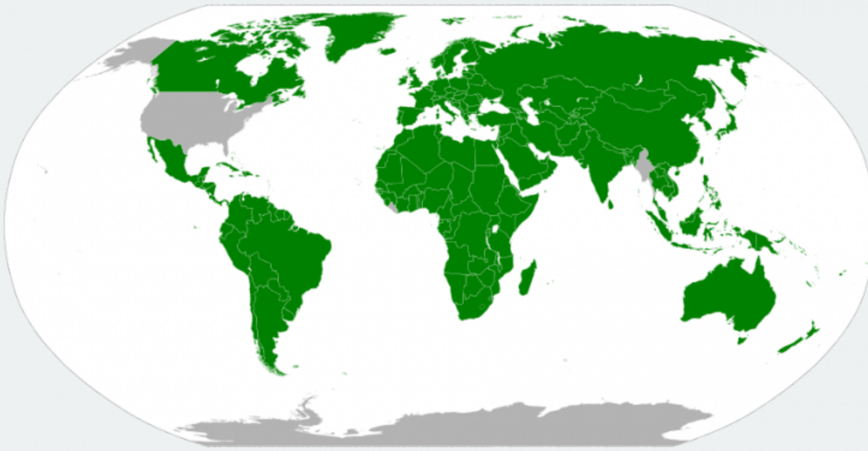

Figure 3.1.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great". Table 3.1.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 3.1.2: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\text{Gm}) = 10^9 \text{ m}$
mega	M	1,000,000	1 megameter $(\text{Mm}) = 10^6 \text{ m}$
kilo	k	1,000	1 kilometer $(\text{km}) = 1,000 \text{ m}$
hecto	h	100	1 hectometer $(\text{hm}) = 100 \text{ m}$
deka	da	10	1 dekameter $(\text{dam}) = 10 \text{ m}$
		1	1 meter (m)
deci	d	1/10	1 decimeter $(\text{dm}) = 0.1 \text{ m}$
centi	c	1/100	1 centimeter $(\text{cm}) = 0.01 \text{ m}$
milli	m	1/1,000	1 millimeter $(\text{mm}) = 0.001 \text{ m}$
micro	μ	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6} \text{ m}$
nano	n	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9} \text{ m}$
pico	p	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12} \text{ m}$

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. We use "m" for meter and not "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So we would write 3.5 milliliters as 3.5 mL.

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs 6.5 kg, this is easier than saying it weighs 6500 g or 0.65 dag. All three are correct, but the kg units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

✓ Example 3.1.1: Unit Abbreviations

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

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

Solutions

Solutions to Example 2.5.1

	Explanation	Answer
a	The prefix kilo means “1,000 ×,” so 1 kL equals 1,000 L.	kL
b	The prefix micro implies 1/1,000,000th of a unit, so 1 µs equals 0.000001 s.	µs
c	The prefix deci means 1/10th, so 1 dm equals 0.1 m.	dm
d	The prefix nano means 1/1000000000, so a nanogram is equal to 0.000000001 g.	ng

? Exercise 3.1.1

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

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

- **2.5: The Basic Units of Measurement** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

3.2: Metric Conversion Factors

Maps take some time to build because we have to find or write matching materials. LibreTexts POV is that it is best to make available pages that we have finished rather than wait till the entire project is complete .This map is not completely finished, some pages are missing but we are workin' on it. . . If you know of some open materials which could be used here let us know info@libretexts.org (Public Domain ; Public Domain Pictures)

3.3: Metric-Metric Conversions

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3.4: Metric-English Conversions

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3.5: Volume by Calculation

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3.6: Volume by Displacement

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3.7: The Density Concept

Learning Objectives

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

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

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

or just

$$\rho = \frac{m}{V} \tag{3.7.1}$$

Based on this equation, it's clear that density can, and does, vary from element to element and substance to substance due to differences in the relationship of mass and volume. Pure water, for example, has a density of 0.998 g/cm^3 at 25° C. The average densities of some common substances are in Table $3.7.1$. 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 $3.7.1$: Densities of Common Substances

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

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

Example $3.7.1$: 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 $3.7.1$:

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

Exercise $3.7.1$

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

Answer a

$0.274\, kg/L$

Answer b

$1.049\, kg/L$

Density as a Conversion Factor

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

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

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

and

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

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

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

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

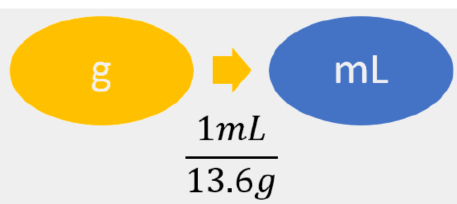
> *Density can be used as a conversion factor between mass and volume.*

Example 3.7.2: Mercury Thermometer Steps for Problem Solving

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

Solution

Solution to Example 2.9.2

Steps for Problem Solving	**Unit Conversion**
Identify the "given" information and what the problem is asking you to "find."	Given: 0.750 g Find: mL
List other known quantities.	13.6 g/mL (density of mercury)
Prepare a concept map.	
Calculate.	$0.750\ \cancel{\text{g}} \times \frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}} = 0.055147\ldots\ \text{mL} \approx 0.0551\ \text{mL}$ We have limited the final answer to three significant figures.

> **?** Exercise 3.7.2
>
> What is the volume of 100.0 g of air if its density is 1.3 g/L?
>
> ***Answer***
>
> $77\ L$

Summary

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

Contributions & Attributions

3.8: Temperature

Learning Objective

- Learn about the various temperature scales that are commonly used in chemistry.
- Define density and use it as a conversion factor.

There are other units in chemistry that are important, and we will cover others over the course of the entire book. One of the fundamental quantities in science is temperature. Temperature is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees, and there are several temperature scales in use. In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as "degrees Fahrenheit"). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

Science also uses other scales to express temperature. The Celsius scale (symbolized by °C and spoken as "degrees Celsius") is a temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water; the scale is divided into 100 divisions between these two landmarks and extended higher and lower. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}\mathrm{C} = (^{\circ}\mathrm{F} - 32) \times \tfrac{5}{9} \tag{3.8.1}$$

$$^{\circ}\mathrm{F} = \left(^{\circ}\mathrm{C} \times \tfrac{9}{5}\right) + 32 \tag{3.8.2}$$

Example 3.8.1: Conversions

a. What is 98.6 °F in degrees Celsius?
b. What is 25.0 °C in degrees Fahrenheit?

Solution

a. Using Equation 3.8.1, we have

$$\begin{aligned} ^{\circ}C &= (98.6 - 32) \times \frac{5}{9} \\ &= 66.6 \times \frac{5}{9} \\ &= 37.0^{\circ}C \end{aligned}$$

b. Using Equation 3.8.2, we have

$$\begin{aligned} ^{\circ}F &= \left(25.0 \times \frac{9}{5}\right) + 32 \\ &= 45.0 + 32 \\ &= 77.0^{\circ}F \end{aligned}$$

Exercise 3.8.1

a. Convert 0 °F to degrees Celsius.
b. Convert 212 °C to degrees Fahrenheit.

Answer a

−17.8 °C

Answer b

414 °F

The fundamental unit of temperature (another fundamental unit of science, bringing us to four) in SI is the kelvin (K). The Kelvin temperature scale (note that the name of the scale capitalizes the word *Kelvin*, but the unit itself is lowercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$K = {}^{\circ}C + 273.15$$

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as "two hundred ninety-five kelvins" and not "two hundred ninety-five degrees Kelvin."

The reason that the Kelvin scale is defined this way is because there exists a minimum possible temperature called absolute zero (zero kelvins). The Kelvin temperature scale is set so that 0 K is absolute zero, and temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

Example $3.8.2$: Room Temperature

If normal room temperature is 72.0°F, what is room temperature in degrees Celsius and kelvin?

Solution

First, we use the formula to determine the temperature in degrees Celsius:

$$\begin{aligned} {}^{\circ}C &= (72.0 - 32) \times \frac{5}{9} \\ &= 40.0 \times \frac{5}{9} \\ &= 22.2^{\circ}C \end{aligned}$$

Then we use the appropriate formula above to determine the temperature in the Kelvin scale:

$$\begin{aligned} K &= 22.2^{\circ}C + 273.15 \\ &= 295.4K \end{aligned}$$

So, room temperature is about 295 K.

Exercise $3.8.2$

What is 98.6°F on the Kelvin scale?

Answer

310.2 K

Figure $3.8.1$ compares the three temperature scales. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. The other two countries are Liberia and Myanmar (formerly Burma). People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.

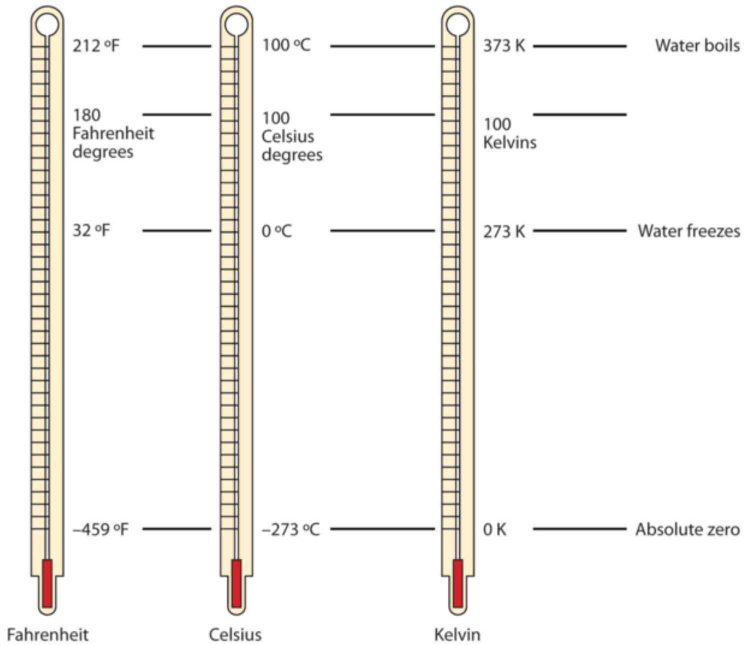


Figure 3.8.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales.

Density

Density is a physical property that is defined as a substance's mass divided by its volume:

$$density = \frac{mass}{volume} \Rightarrow d = \frac{m}{v}$$

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances are listed in Table $3.8.1$.

Table 3.8.1: Densities of Some Common Substances

Substance	Density (g/mL or g/cm^3)
water	1.0
gold	19.3
mercury	13.6
air	0.0012
cork	0.22–0.26
aluminum	2.7
iron	7.87

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table $3.8.1$), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

$$7.88\ \cancel{cm^3} \times \frac{2.7\ g}{\cancel{cm^3}} = 21\ g \text{ of aluminium}$$

where we have limited our answer to two significant figures.

Example 3.8.3: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 3.8.1 "Densities of Some Common Substances" as a conversion factor to go from volume to mass:

$$44.6\ \cancel{mL} \times \frac{13.6\,g}{\cancel{mL}} = 607\,g$$

The mass of the mercury is 607 g.

Exercise 3.8.3

What is the mass of 25.0 cm^3 of iron?

Answer

197 g

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3\,g/mL = \frac{19.3\,g}{mL}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator:

$$d = 19.3\,g/mL = \frac{19.3\,g}{mL}$$

That is, the density value tells us that we have 19.3 grams for every 1 milliliter of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched—that is, we must take the *reciprocal* of the density. In so doing, we move not only the units, but also the numbers:

$$\frac{1}{d} = \frac{1\,mL}{19.3\,g}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9\ \cancel{g} \times \frac{1\,mL}{19.3\ \cancel{g}} = 2.38\,mL$$

Note how the mass units cancel, leaving the volume unit, which is what we are looking for.

Example 3.8.4: Calculating Volume from Density

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm^3, what is the volume of the cork?

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find:

$$\frac{1}{d} = \frac{1\,cm^3}{0.22\,g}$$

We can use this expression as the conversion factor. So

$$3.78 \cancel{g} \times \frac{1\,cm^3}{0.22\,\cancel{g}} = 17.2\,cm^3$$

Exercise 3.8.4

What is the volume of 3.78 g of gold?

Answer

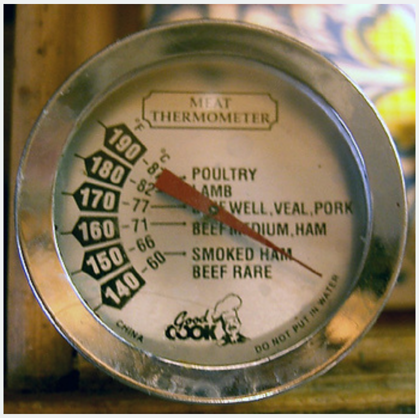
0.196 cm^3

Care must be used with density as a conversion factor. Make sure the mass units are the same, or the volume units are the same, before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Food and Drink Application: Cooking Temperatures

Because degrees Fahrenheit is the common temperature scale in the United States, kitchen appliances, such as ovens, are calibrated in that scale. A cool oven may be only 150°F, while a cake may be baked at 350°F and a chicken roasted at 400°F. The broil setting on many ovens is 500°F, which is typically the highest temperature setting on a household oven.

People who live at high altitudes, typically 2,000 ft above sea level or higher, are sometimes urged to use slightly different cooking instructions on some products, such as cakes and bread, because water boils at a lower temperature the higher in altitude you go, meaning that foods cook slower. For example, in Cleveland water typically boils at 212°F (100°C), but in Denver, the Mile-High City, water boils at about 200°F (93.3°C), which can significantly lengthen cooking times. Good cooks need to be aware of this.



A meat thermometer with a dial. Notice the markings for Fahrenheit (outer scale) and Celsius (inner scale) temperatures. Recipes for cooking food in an oven can use very different numbers, depending on the country you're in. (CC BY2.0 Bev Sykes)

At the other end is pressure cooking. A pressure cooker is a closed vessel that allows steam to build up additional pressure, which increases the temperature at which water boils. A good pressure cooker can get to temperatures as high as 252°F (122°C); at these temperatures, food cooks much faster than it normally would. Great care must be used with pressure cookers because of the high pressure and high temperature. (When a pressure cooker is used to sterilize medical instruments, it is called an *autoclave*.)

Other countries use the Celsius scale for everyday purposes. Therefore, oven dials in their kitchens are marked in degrees Celsius. It can be confusing for US cooks to use ovens abroad—a 425°F oven in the United States is equivalent to a 220°C oven in other countries. These days, many oven thermometers are marked with both temperature scales.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates the mass and volume of a substance.
- Density can be used to calculate volume from a given mass or mass from a given volume.

3.9: 2.9- Heat and Specific Heat

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. The heat capacity of an object depends 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^oC). 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^oC. 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 $\left(\text{J/g}^\text{o} \text{C} \right)$ or calories per gram per degree $\left(\text{cal/g}^\text{o} \text{C} \right)$ (Table $3.9.1$). This text will use $\text{J/g}^\text{o} \text{C}$ for specific heat.

$$\text{specific heat} = \frac{\text{heat}}{\text{mass} \times \text{cal/g}^\text{o} \text{C}}$$

Notice that water has a very high specific heat compared to most other substances.

Table 3.9.1: Specific Heat Capacities

Substance	Specific Heat Capacity at 25$^\text{o}$C in J/g $^\text{o}$C	Substance	Specific Heat Capacity at 25$^\text{o}$C in J/g $^\text{o}$C
H_2 gas	14.267	steam @ 100$^\text{o}$C	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 @ 0$^\text{o}$C	2.010	concrete	0.880
steam @ 100$^\text{o}$C	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 3.9.1: 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.

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by $1^{\circ}\mathrm{C}$).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1^{\circ}\mathrm{C}$.

3.E: The Metric System (Exercises)

Maps take some time to build because we have to find or write matching materials. LibreTexts POV is that it is best to make available pages that we have finished rather than wait till the entire project is complete .This map is not completely finished, some pages are missing but we are workin' on it. . . If you know of some open materials which could be used here let us know info@libretexts.org (Public Domain ; Public Domain Pictures)

CHAPTER OVERVIEW

4: Matter and Energy

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

- **Homework**

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4.1: Conservation of Energy

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4.2: 3.E - Matter and Energy (Exercises)

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4.3: Physical States of Matter

Learning Objectives

- To describe the solid, liquid and gas phases.

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

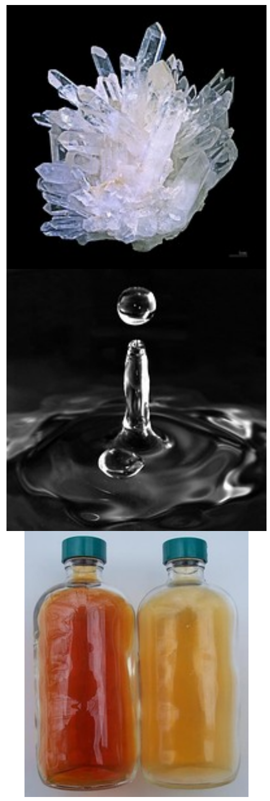


Figure $4.3.1$: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state a given substance exhibits is also a physical property. 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. All substances can exist in any of these three states. Figure $4.3.2$ shows the differences among solids, liquids, and gases at the molecular 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.

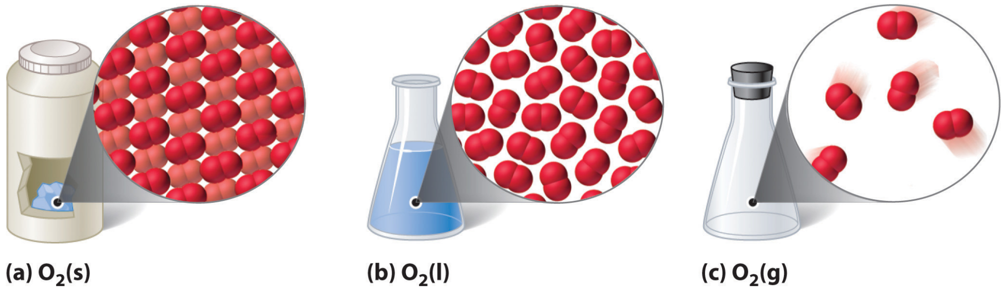
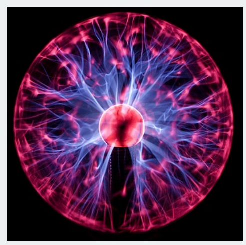


Figure 4.3.2: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 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 intermolecular interactions between the particles. 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 $-39^{o}C$, 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 usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure $4.3.3$. 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.

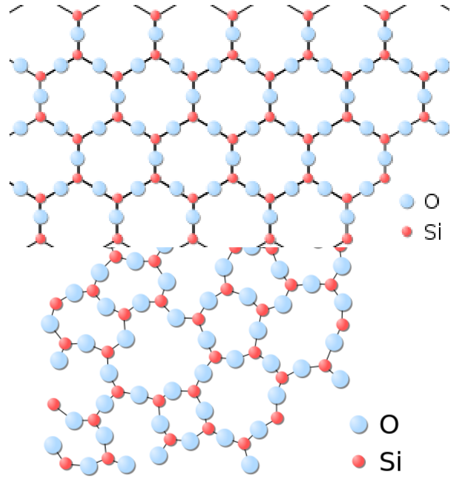


Figure 4.3.3: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in 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 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 4.3.1**: Mercury boiling to become a gas.*

If we heat liquid mercury to its boiling point of $357^{\circ}C$, and 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, which we will consider in more detail elsewhere. 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
- Particles move in random motion with little or no attraction to each other
- Highly compressible

Table 4.3.1: Characteristics of the Three States of Matter

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

Example 4.3.1

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 4.3.1

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

Summary

- Three states of matter exist - solid, liquid, and gas.
- Solids have a definite shape and volume.
- Liquids have a definite volume, but take the shape of the container.
- Gases have no definite shape or volume

Contributors and Attributions

-
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- Henry Agnew (UC Davis)

4.4: Elements, Compounds, and Mixtures

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4.5: Names and Symbols of the Elements

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4.6: Metals, Nonmetals, and Semimetals

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4.7: Compounds and Chemical Formulas

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4.8: Physical and Chemical Properties

> **Learning Objectives**
>
> To separate physical from chemical properties.

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: Pencil (left) and 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 ^oC.	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

Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed.

4.9: Physical and Chemical Changes

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4.10: Conservation of Mass

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4.11: Potential and Kinetic Energy

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CHAPTER OVERVIEW

5: Models of the Atom

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

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5.1: Electron Configuration

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5.2: Quantum Mechanical Model of the Atom

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5.3: Dalton Model of the Atom

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5.5: Rutherford Model of the Atom

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5.6: Atomic Notation

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5.7: Atomic Mass

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5.8: The Wave Nature of Light

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5.9: The Quantum Concept

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5.10: Bohr Model of the Atom

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5.E: Models of the Atom (Exercises)

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CHAPTER OVERVIEW

6: The Periodic Table

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by Charles Corwin

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- **Homework**

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6.1: Ionic Charges

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6.2: Classification of Elements

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6.4: Groups and Periods of Elements

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6.6: Properties of Elements

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6.7: Blocks of Elements

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6.8: Valence Electrons

6.9: Electron Dot Formulas

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6.10: Ionization Energy

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6.E: The Periodic Table

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CHAPTER OVERVIEW

7: Language of Chemistry

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7.1: Classification of Compounds

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7.3: Polyatomic Ions

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7.4: Writing Chemical Formulas

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7.5: Binary Ionic Compounds

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7.6: Ternary Ionic Compounds

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7.7: Binary Molecular Compounds

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7.8: Binary Acids

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7.9: Ternary Oxyacids

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7.E: Exercises

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CHAPTER OVERVIEW

8: Chemical Reactions

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

- **Homework**

Template:HideTOC

8.1: Double-Replacement Reactions

8.2: Neutralization Reactions

8.3: Evidence for Chemical Reactions

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Edit page

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Tips:

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8.4: Writing Chemical Equations

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8.5: Balancing Chemical Equations

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8.6: Classifying Chemical Reactions

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8.7: Combination Reactions

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8.8: Decomposition Reactions

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8.9: The Activity Series Concept

8.10: Single-Replacement Reactions

8.11: Solubility Rules

CHAPTER OVERVIEW

9: The Mole Concept

9.1: Avogadro's Number

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9.2: Mole Calculations I- atom conversions

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9.3: Molar Mass

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9.4: Mole Calculations II- mass conversions

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9.5: Molar Volume

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9.6: Mole Calculations III- Gas volume conversions

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9.7: Percent Composition

9.8: Empirical Formula

9.9: Molecular Formula

9.10: Section 10

CHAPTER OVERVIEW

10: Chemical Equation Calculations

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

- **Homework**

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10.1: Interpreting a Chemical Equation

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10.2: Mole-Mole Relationships

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10.3: Types of Stoichiometry Problems

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10.4: Mass-Mass Problems

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10.5: Mass-Volume Problems

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10.6: Volume-Volume Problems

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10.7: The Limiting Reactant Concept

10.8: Limiting Reactant Problems

10.9: Percent Yield

10.E: Chemical Equation Calculations (Exercises)

CHAPTER OVERVIEW

11: The Gaseous State

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

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11.1: Properties of Gases

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11.2: Atmospheric Pressure

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11.3: Variables Affecting Gas Pressure

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11.4: Boyle's Law- Pressure-Volume Relationships

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11.5: Charles's Law- Volume-Temperature Relationships

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11.6: Gay-Lussac's Law- Pressure-Temperature Relationships

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11.7: Combined Gas Law

11.8: The Vapor Pressure Concept

11.9: Dalton's Law of Partial Pressures

11.10: Ideal Gas Behavior

11.11: Ideal Gas Law

CHAPTER OVERVIEW

12: Liquids and Solids

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

Template:HideTOC

12.1: Properties of Liquids

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12.2: The Intermolecular Bond Concept

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12.3: Vapor Pressure, Boiling Point, Viscosity, Surface Tension

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12.4: Properties of Solids

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12.5: Crystalline Solids

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12.6: Changes of Physical State

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12.7: Structure of Water

12.8: Physical Properties of Water

12.9: Chemical Properties of Water

12.10: Hydrates

13: Chemical Bonding

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

Template:HideTOC

13.1: The Chemical Bond Concept

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13.2: Ionic Bonds

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13.3: Covalent Bonds

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13.4: Electron Dot Formulas of Molecules

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13.5: Electron Dot Formulas of Polyatomic Ions

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13.6: Polar Covalent Bonds

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13.7: Nonpolar Covalent Bonds

13.8: Coordinate Covalent Bonds

13.9: Hydrogen Bonds

13.10: Shapes of Molecules

CHAPTER OVERVIEW

14: Solutions

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

Template:HideTOC

14.1: Gases in Solution

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14.2: Liquids in Solution

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14.3: Solids in Solution

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14.4: The Dissolving Process

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14.5: Rate of Dissolving

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14.6: Solubility and Temperature

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14.7: Unsaturated, Saturated, and Supersaturated Solutions

14.8: Mass/Mass Percent Concentration

14.9: Molar Concentration

14.10: Dilution of a Solution

14.11: Solution Stoichiometry

CHAPTER OVERVIEW

15: Acids and Bases

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

I II III IV V VI VII VIII IX X XI XII XIII XIV XV XVI XVII XVIII XIX XX

- **Homework**

Template:HideTOC

15.1: Properties of Acids and Bases

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15.2: Arrhenius Acids and Bases

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15.3: Brønsted-Lowry Acids and Bases

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15.4: Acid-Base Indicators

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15.5: Acid-Base Titrations

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15.6: Acid-Base Standardization

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15.7: Ionization of Water

15.8: The pH Concept

15.9: Advanced pH Calculations

15.10: Strong and Weak Electrolytes

15.11: Net Ionic Equations

CHAPTER OVERVIEW

16: Chemical Equilibrium

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16.1: Collision Theory

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16.2: Energy Profiles of Chemical Reactions

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16.3: The Chemical Equilibrium Concept

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16.4: General Equilibrium Constant - K eq Keq

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16.5: Equilibria Shifts for Gases

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16.6: Ionization Equilibrium Constant - K i Ki

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16.7: Equilibria Shifts for Weak Acids and Bases

16.8: Solubility Product Equilibrium Constant - K sp Ksp

16.9: Equilibria Shifts for Slightly Soluble Compounds

CHAPTER OVERVIEW

17: Oxidation and Reduction

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

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17.1: Oxidation Numbers

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17.2: Oxidation-Reduction Reactions

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17.3: Balancing Redox Equations - Oxidation Number Method

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17.4: Balancing Redox Equations - Half-Reaction Method

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17.5: Predicting Spontaneous Redox Reactions

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17.6: Voltaic Cells

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17.7: Electrolytic Cells

CHAPTER OVERVIEW

18: Nuclear Chemistry

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by Charles Corwin

Template:HideTOC

18.1: Natural Radioactivity

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18.2: Nuclear Equations

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18.3: Radioactive Decay Series

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18.4: Radioactive Half-Life

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18.5: Applications of Radionuclides

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18.6: Induced Radioactivity

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18.7: Nuclear Fission

18.8: Nuclear Fusion

CHAPTER OVERVIEW

19: Organic Chemistry

An introductory chemistry Libretexts Textmap organized around the textbook

Introductory Chemistry: Concepts & Critical Thinking

by Charles Corwin

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19.1: Hydrocarbons

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19.2: Alkanes

19.3: Alkenes and Alkynes

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19.4: Arenes

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19.5: Hydrocarbon Derivatives

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19.6: Organic Halides

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19.7: Alcohols, Phenols, and Ethers

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19.8: Amines

19.9: Aldehydes and Ketones

19.10: Carboxylic Acids, Esters, and Amides

CHAPTER OVERVIEW

20: Biochemistry

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by Charles Corwin

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20.1: 20.1- Biological Compounds

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20.2: Proteins

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20.3: Enzymes

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20.4: Carbohydrates

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20.5: Lipids

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20.6: Nucleic Acids

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Glossary

Acetyl-CoA | A molecule that participates in many biochemical reactions in protein, carbohydrate and lipid metabolism. Its main function is to deliver the acetyl group to the citric acid cycle (Krebs cycle) to be oxidized for energy production.

addition polymerization | A reaction in which monomers add to one another to produce a polymeric product that contains all the atoms of the starting monomers.

addition reactions | A reaction in which substituent groups join to hydrocarbon molecules at points of unsaturation—the double or triple bonds.

alcohol | An organic compound with an OH functional group on an aliphatic carbon atom.

aldehyde | An organic compound with a carbonyl functional group that has an hydrogen atom attached and either a hydrocarbon group or a second hydrogen atom.

alkaloid | A nitrogen-containing organic compound obtained from plants that has physiological properties.

alkanes (or saturated hydrocarbons) | A hydrocarbon with only carbon-to-carbon single bonds and existing as a continuous chain of carbon atoms also bonded to hydrogen atoms

alkenes | A hydrocarbon with one or more carbon–carbon double bonds.

alkyl group | A hydrocarbon group derived from an alkane by removal of a hydrogen atom.

alkyl halide (or haloalkane) | A compound resulting from the replacement of a hydrogen atom of an alkane with a halogen atom.

Alkynes | A hydrocarbon with a carbon–carbon triple bond.

alloy | A solid solution of a metal with other substances dissolved in it.

alpha particle | A type of radioactive emission that is equivalent to a helium atom nucleus.

amide | An organic compound with a carbonyl group joined to a nitrogen atom from ammonia or an amine.

amine | An organic compound derived from ammonia by the replacement of one, two, or three of the hydrogens atoms by alkyl or aryl groups.

amino group | An NH_2 unit.

amorphous | A solid with no regular structure.

amphiprotic | A substance that can either donate or accept a proton, depending on the circumstances.

amylopectin | A branched polymer of glucose units found in starch.

anabolism | Metabolic reactions in which molecules are synthesized.

anaerobic metabolism | A biochemical process that takes place in the absence of oxygen.

anions | A negatively charged ion.

anomeric carbon | The carbon atom that was the carbonyl carbon atom in the straight-chain form of a monosaccharide.

anticodon | A set of three nucleotides on the tRNA that is complementary to, and pairs with, the codon on the mRNA.

anticodon | A set of three nucleotides on the tRNA that is complementary to, and pairs with, the codon on the mRNA.

antioxidants | A substance in foods that acts as a reducing agent.

antioxidants | A compound that prevents oxidation.

aromatic compound | Any compound that contains a benzene ring or has certain benzene-like properties.

aromatic hydrocarbons | A hydrocarbon with a benzene-like structure.

aryl group | A group derived from an aromatic hydrocarbon by the removal of a hydrogen atom.

atomic bomb | A weapon that depends on a nuclear chain reaction to generate immense forces.

atomic mass | A weighted average of the masses of all the element's naturally occurring isotopes.

atomic mass unit | One-twelfth the mass of a ^{12}C atom.

atomic radius | The approximate size of an atom.

autoionization of water | The process by which water ionizes into hydronium ions and hydroxide ions as it acts as an acid and a base.

Avogadro's number | The value 6.022×10^{23}.

balanced | A property of a chemical equation when there are the same number of atoms of each element in the reactants and products.

base | A compound that increases the concentration of hydroxide ion ($OH^−$) in aqueous solution.

Base (or basic) units | A fundamental unit of SI.

beta particle | A type of radioactive emission that is equivalent to an electron.

beta particle | A type of radioactive emission that is equivalent to an electron.

Bilayers | A double layer of lipids arranged so that nonpolar tails are found between an inner surface and outer surface consisting of hydrophilic heads.

Bilayers | A double layer of lipids arranged so that nonpolar tails are found between an inner surface and outer surface consisting of hydrophilic heads.

Bile | The yellowish green liquid produced in the liver.

Bile | The yellowish green liquid produced in the liver.

biochemistry | The chemistry of molecules found in living organisms.

biochemistry | The chemistry of molecules found in living organisms.

boiling point | The temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid).

boiling point | The temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid).

boiling point elevation | The raising of the boiling point of a solution versus the pure solvent.

boiling point elevation | The raising of the boiling point of a solution versus the pure solvent.

bond length | The distance between two nuclei in a covalent bond.

bond length | The distance between two nuclei in a covalent bond.

Boyle's law | The gas law that relates pressure and volume.

Boyle's law | The gas law that relates pressure and volume.

Brønsted-Lowry base | A compound that accepts a hydrogen ion (H^+) in a reaction; a proton acceptor.

Brønsted-Lowry base | A compound that accepts a hydrogen ion (H^+) in a reaction; a proton acceptor.

buffer | A solution that resists dramatic changes in pH.

buffer | A solution that resists dramatic changes in pH.

calorie | A unit of energy widely used in the health professions and everyday life.

calorie | A unit of energy widely used in the health professions and everyday life.

capacity | The amount of strong acid or base a buffer can counteract.

capacity | The amount of strong acid or base a buffer can counteract.

carbohydrates | A compound composed of carbon, hydrogen, and oxygen atoms that is a polyhydroxy aldehyde or ketone or a compound that can be broken down to form such a compound. It is one of the three main components of the human diet.

carbohydrates | A compound composed of carbon, hydrogen, and oxygen atoms that is a polyhydroxy aldehyde or ketone or a compound that can be broken down to form such a compound. It is one of the three main components of the human diet.

carbonyl group | A compound with an carbon-to-oxygen double bond.

carbonyl group | A compound with an carbon-to-oxygen double bond.

carboxyl group | A functional group that contains a carbon–oxygen bond and an OH group also attached to the same carbon atom.

carboxyl group | A functional group that contains a carbon–oxygen bond and an OH group also attached to the same carbon atom.

carboxylic acids | An organic compound that has a carboxyl functional group.

carboxylic acids | An organic compound that has a carboxyl functional group.

cations | A positively charged ion.

cations | A positively charged ion.

cerebrosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and galactose or glucose.

cerebrosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and galactose or glucose.

chain reaction | An exponential growth in a process.

chain reaction | An exponential growth in a process.

Charles's law | The gas law that relates volume and absolute temperature.

Charles's law | The gas law that relates volume and absolute temperature.

chemical bond | A very strong attraction between two atoms.

chemical bond | A very strong attraction between two atoms.

chemical equilibrium (or equilibrium) | The condition in which the extent of a chemical reaction does not change any further.

chemical equilibrium (or equilibrium) | The condition in which the extent of a chemical reaction does not change any further.

chemical formula | A concise list of the elements in a compound and the ratios of these elements.

chemical formula | A concise list of the elements in a compound and the ratios of these elements.

Chemical properties | A characteristic that describes how matter changes its chemical structure or composition.

Chemical properties | A characteristic that describes how matter changes its chemical structure or composition.

chemical reaction | A representation of a chemical change.

chemical reaction | A representation of a chemical change.

chemical symbol | A one- or two-letter abbreviation for an element.

chemical symbol | A one- or two-letter abbreviation for an element.

chiral carbon | A carbon atom that has four different groups attached to it.

chiral carbon | A carbon atom that has four different groups attached to it.

Cholesterol | A steroid that is found in mammals.

Cholesterol | A steroid that is found in mammals.

cis-trans isomers (or geometric isomers) | Isomers that have different configurations because of the presence of a rigid structure such as a double bond or ring.

cis-trans isomers (or geometric isomers) | Isomers that have different configurations because of the presence of a rigid structure such as a double bond or ring.

citric acid cycle (or Krebs cycle or tricarboxylic acid [TCA] cycle) | A cyclic sequence of reactions that brings about the oxidation of a two-C unit to carbon dioxide and water.

citric acid cycle (or Krebs cycle or tricarboxylic acid [TCA] cycle) | A cyclic sequence of reactions that brings about the oxidation of a two-C unit to carbon dioxide and water.

codon | A set of three nucleotides on the mRNA that specifies a particular amino acid.

codon | A set of three nucleotides on the mRNA that specifies a particular amino acid.

coefficient | A number that gives the number of molecules of a substance in a balanced chemical equation.

coefficient | A number that gives the number of molecules of a substance in a balanced chemical equation.

coenzymes | A cofactor that is an organic molecule.

coenzymes | A cofactor that is an organic molecule.

colligative properties | A characteristic of solutions that depends only on the number of dissolved particles.

colligative properties | A characteristic of solutions that depends only on the number of dissolved particles.

combination (composition) reaction | A chemical reaction that makes a single substance from two or more reactants.

combination (composition) reaction | A chemical reaction that makes a single substance from two or more reactants.

combined gas law | The gas law that relates pressure, volume, and absolute temperature.

combined gas law | The gas law that relates pressure, volume, and absolute temperature.

combustion reaction | A chemical reaction accompanied by the production of light and/or heat, typically a vigorous reaction because of combination with oxygen.

combustion reaction | A chemical reaction accompanied by the production of light and/or heat, typically a vigorous reaction because of combination with oxygen.

combustion reaction | A chemical reaction in which a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction.

combustion reaction | A chemical reaction in which a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction.

competitive inhibitor | A compound that resembles a particular substrate and competes with the substrate for binding at the active site of an enzyme to slow the rate of the reaction.

competitive inhibitor | A compound that resembles a particular substrate and competes with the substrate for binding at the active site of an enzyme to slow the rate of the reaction.

complementary bases | Specific base pairings in the DNA double helix.

complementary bases | Specific base pairings in the DNA double helix.

compound | A substance that can be broken down into chemically simpler components.

compound | A substance that can be broken down into chemically simpler components.

concentration | How much solute is dissolved in a certain amount of solvent.

concentration | How much solute is dissolved in a certain amount of solvent.

condensed structural formulas | An organic chemical formula that shows the hydrogen atoms (or other atoms or groups) right next to the carbon atoms to which they are attached.

condensed structural formulas | An organic chemical formula that shows the hydrogen atoms (or other atoms or groups) right next to the carbon atoms to which they are attached.

conversion factor | A fraction that has equivalent quantities in the numerator and the denominator but expressed in different units.

conversion factor | A fraction that has equivalent quantities in the numerator and the denominator but expressed in different units.

core electrons | An electron in a lower-numbered shell of an atom.

core electrons | An electron in a lower-numbered shell of an atom.

covalent network bonding | A type of interaction in which all the atoms in a sample are covalently bonded to other atoms.

covalent network bonding | A type of interaction in which all the atoms in a sample are covalently bonded to other atoms.

curie (Ci) | A unit of radioactivity equal to 3.7×10^{10} decays per second.

curie (Ci) | A unit of radioactivity equal to 3.7×10^{10} decays per second.

cyclic hydrocarbons | A hydrocarbon with a ring of carbon atoms.

cyclic hydrocarbons | A hydrocarbon with a ring of carbon atoms.

cycloalkanes | A cyclic hydrocarbon with only single bonds.

cycloalkanes | A cyclic hydrocarbon with only single bonds.

cytochromes | A protein that contains an iron porphyrin in which iron can alternate between Fe(II) and Fe(III).

cytochromes | A protein that contains an iron porphyrin in which iron can alternate between Fe(II) and Fe(III).

cytoplasm | Everything between the cell membrane and the nuclear membrane.

cytoplasm | Everything between the cell membrane and the nuclear membrane.

decomposition reaction | A chemical reaction in which a single substance is converted into two or more products.

decomposition reaction | A chemical reaction in which a single substance is converted into two or more products.

Denaturation | Any change in the three-dimensional structure of a macromolecule that renders it incapable of performing its assigned function.

Denaturation | Any change in the three-dimensional structure of a macromolecule that renders it incapable of performing its assigned function.

Density | The mass of an object divided by its volume.

Density | The mass of an object divided by its volume.

Derived units | A combinations of the SI base units.

Derived units | A combinations of the SI base units.

diatomic molecules | A two-atom grouping that behaves as a single chemical entity.

diatomic molecules | A two-atom grouping that behaves as a single chemical entity.

digestion | The breakdown of food molecules by hydrolysis reactions into the individual monomer units in the mouth, stomach, and small intestine.

digestion | The breakdown of food molecules by hydrolysis reactions into the individual monomer units in the mouth, stomach, and small intestine.

dipole-dipole interaction | An attraction between polar molecules.

dipole-dipole interaction | An attraction between polar molecules.

Dispersion forces | A force caused by the instantaneous imbalance of electrons about a molecule.

Dispersion forces | A force caused by the instantaneous imbalance of electrons about a molecule.

dispersion forces (or London forces) | A force caused by the instantaneous imbalance of electrons about a molecule.

dispersion forces (or London forces) | A force caused by the instantaneous imbalance of electrons about a molecule.

dissociation | The process of cations and anions of an ionic solute separating when the solute dissolves.

dissociation | The process of cations and anions of an ionic solute separating when the solute dissolves.

Disulfide linkages | A covalent bond that forms by the oxidation and linkage of two sulfur atoms from the side chains of two cysteine residues.

Disulfide linkages | A covalent bond that forms by the oxidation and linkage of two sulfur atoms from the side chains of two cysteine residues.

double bond | Two pairs of electrons being shared by two atoms in a molecule.

double bond | Two pairs of electrons being shared by two atoms in a molecule.

electron | A subatomic particle with a negative electric charge.

electron | A subatomic particle with a negative electric charge.

electron configuration | A shorthand description of the arrangement of electrons in an atom.

electron configuration | A shorthand description of the arrangement of electrons in an atom.

electron transport chain (or respiratory chain) | An organized sequence of oxidation-reduction reactions that ultimately transports electrons to oxygen, reducing it to water.

electron transport chain (or respiratory chain) | An organized sequence of oxidation-reduction reactions that ultimately transports electrons to oxygen, reducing it to water.

electronegativity | A relative measure of how strongly an atom attracts electrons when it forms a covalent bond.

electronegativity | A relative measure of how strongly an atom attracts electrons when it forms a covalent bond.

emulsion | A dispersion of two liquids that do not normally mix.

emulsion | A dispersion of two liquids that do not normally mix.

enantiomers | Stereoisomers that are nonsuperimposable mirror images of each other.

enantiomers | Stereoisomers that are nonsuperimposable mirror images of each other.

endothermic | A process that absorbs energy.

endothermic | A process that absorbs energy.

Energy | The ability to do work.

Energy | The ability to do work.

equivalents (Eq) | One mole of charge (either positive or negative).

equivalents (Eq) | One mole of charge (either positive or negative).

essential amino acids | An amino acid that must be obtained from the diet because it cannot be synthesized in sufficient quantities by the body.

essential amino acids | An amino acid that must be obtained from the diet because it cannot be synthesized in sufficient quantities by the body.

essential fatty acids | A fatty acid that must be obtained from the diet because it cannot be synthesized by the human body.

essential fatty acids | A fatty acid that must be obtained from the diet because it cannot be synthesized by the human body.

ester | An organic compound derived from a carboxylic acid and an alcohol in which the OH of the acid is replaced by an OR group.

ester | An organic compound derived from a carboxylic acid and an alcohol in which the OH of the acid is replaced by an OR group.

esterification | The formation of an ester from a carboxylic acid and an alcohol.

esterification | The formation of an ester from a carboxylic acid and an alcohol.

ether | An organic compound that has an oxygen atom between two hydrocarbon groups.

ether | An organic compound that has an oxygen atom between two hydrocarbon groups.

Exact numbers | A number that is defined or counted.

Exact numbers | A number that is defined or counted.

exothermic | A process that gives off energy.

exothermic | A process that gives off energy.

fats | A compound, composed largely of hydrocarbon chains, that supplies energy for the body.

fats | A compound, composed largely of hydrocarbon chains, that supplies energy for the body.

Feedback inhibition | A normal biochemical process that makes use of noncompetitive inhibitors to control some enzymatic activity.

Feedback inhibition | A normal biochemical process that makes use of noncompetitive inhibitors to control some enzymatic activity.

fibrous proteins | A protein that is elongated or fiberlike and insoluble in water.

fibrous proteins | A protein that is elongated or fiberlike and insoluble in water.

formula mass | The sum of the masses of the elements in the formula of an ionic compound.

formula mass | The sum of the masses of the elements in the formula of an ionic compound.

formula unit | A set of oppositely charged ions that compose an ionic compound.

formula unit | A set of oppositely charged ions that compose an ionic compound.

freezing point depression | The lowering of the freezing point of a solution versus the pure solvent.

freezing point depression | The lowering of the freezing point of a solution versus the pure solvent.

functional group | A structural arrangement of atoms and/or bonds that imparts a wide range of important properties to organic compounds.

functional group | A structural arrangement of atoms and/or bonds that imparts a wide range of important properties to organic compounds.

functional groups | A specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to a molecule.

functional groups | A specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to a molecule.

Fusion | A nuclear process in which small nuclei are combined into larger nuclei, releasing energy.

Fusion | A nuclear process in which small nuclei are combined into larger nuclei, releasing energy.

galactosemia | A genetic disease caused by the absence of one of the enzymes needed to convert galactose to glucose.

galactosemia | A genetic disease caused by the absence of one of the enzymes needed to convert galactose to glucose.

gamma rays | A type of radioactive emission that is a very energetic form of electromagnetic radiation.

gamma rays | A type of radioactive emission that is a very energetic form of electromagnetic radiation.

gangliosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and a complex oligosaccharide.

gangliosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and a complex oligosaccharide.

gas law | A simple mathematical formula that relates two or more properties of a gas.

gas law | A simple mathematical formula that relates two or more properties of a gas.

Gastric juice | A mixture of water, inorganic ions, hydrochloric acid, and various enzymes and proteins found in the stomach.

Gastric juice | A mixture of water, inorganic ions, hydrochloric acid, and various enzymes and proteins found in the stomach.

Geiger counter | An electrical device that detects radioactivity.

Geiger counter | An electrical device that detects radioactivity.

genes | The basic unit of heredity.

genes | The basic unit of heredity.

genetic code | The identification of each group of three nucleotides and its particular amino acid.

genetic code | The identification of each group of three nucleotides and its particular amino acid.

genetic diseases | A hereditary condition caused by an altered DNA sequence.

genetic diseases | A hereditary condition caused by an altered DNA sequence.

Globular proteins | A protein that is generally spherical in structure and soluble in water.

Globular proteins | A protein that is generally spherical in structure and soluble in water.

glycols | An alcohol with two OH functional groups.

glycols | An alcohol with two OH functional groups.

glycolysis | The metabolic pathway in which glucose is broken down to two molecules of pyruvate with the corresponding production of ATP.

glycolysis | The metabolic pathway in which glucose is broken down to two molecules of pyruvate with the corresponding production of ATP.

glycosidic linkage | The carbon–oxygen-carbon linkage between monosaccharide units in more complex carbohydrates, such as disaccharides or polysaccharides.

glycosidic linkage | The carbon–oxygen-carbon linkage between monosaccharide units in more complex carbohydrates, such as disaccharides or polysaccharides.

groups (or families) | A column of elements on the periodic table.

groups (or families) | A column of elements on the periodic table.

half reactions | A chemical reaction that shows only oxidation or reduction.

half reactions | A chemical reaction that shows only oxidation or reduction.

half-life | The amount of time it takes for one-half of a radioactive isotope to decay.

half-life | The amount of time it takes for one-half of a radioactive isotope to decay.

halogenated hydrocarbons | A hydrocarbon in which one or more hydrogen atoms has been replaced by a halogen atom.

halogenated hydrocarbons | A hydrocarbon in which one or more hydrogen atoms has been replaced by a halogen atom.

halogenation | A reaction in which a halogen reacts at a carbon-to-carbon double or triple bond to add halogen atoms to carbon atoms.

halogenation | A reaction in which a halogen reacts at a carbon-to-carbon double or triple bond to add halogen atoms to carbon atoms.

heat | The transfer of energy from one part of the universe to another due to temperature differences.

heat | The transfer of energy from one part of the universe to another due to temperature differences.

heat of vaporization | The amount of heat per gram or per mole required for a phase change that occurs at the boiling point.

heat of vaporization | The amount of heat per gram or per mole required for a phase change that occurs at the boiling point.

heterocyclic compounds | A cyclic compound in which one or more atoms in the ring is an element other than a carbon atom.

heterocyclic compounds | A cyclic compound in which one or more atoms in the ring is an element other than a carbon atom.

homogeneous mixtures (or solutions) | A mixture that acts as a single substance so that it is not obvious that two or more substances are present.

homogeneous mixtures (or solutions) | A mixture that acts as a single substance so that it is not obvious that two or more substances are present.

homologous series | Any family of compounds in which adjacent members differ from each other by a definite factor.

homologous series | Any family of compounds in which adjacent members differ from each other by a definite factor.

hydration | Solvation by water molecules.

hydration | Solvation by water molecules.

hydration | The addition of water to a substance; in organic chemistry, the addition of water across the carbon-to-carbon double bond of an alkene or the carbon-to-oxygen double bond of an aldehyde or ketone.

hydration | The addition of water to a substance; in organic chemistry, the addition of water across the carbon-to-carbon double bond of an alkene or the carbon-to-oxygen double bond of an aldehyde or ketone.

hydrogen bonding | A particularly strong type of dipole-dipole interaction caused by a hydrogen atom being bonded to a very electronegative element.

hydrogen bonding | A particularly strong type of dipole-dipole interaction caused by a hydrogen atom being bonded to a very electronegative element.

Hydrogen bonding | Bonding between a highly electronegative oxygen atom or nitrogen atom and a hydrogen atom attached to another oxygen atom or nitrogen atom.

Hydrogen bonding | Bonding between a highly electronegative oxygen atom or nitrogen atom and a hydrogen atom attached to another oxygen atom or nitrogen atom.

hydrogenation | A reaction in which hydrogen gas reacts at a carbon-to-carbon double or triple bond or a carbon-to-oxygen double bond to add hydrogen atoms to carbon atoms.

hydrogenation | A reaction in which hydrogen gas reacts at a carbon-to-carbon double or triple bond or a carbon-to-oxygen double bond to add hydrogen atoms to carbon atoms.

hydrolysis | The reaction of a substance with water.

hydrolysis | The reaction of a substance with water.

ideal gas law | The gas law that relates volume, pressure, temperature, and amount of a gas.

ideal gas law | The gas law that relates volume, pressure, temperature, and amount of a gas.

ideal gas law constant | The constant the appears in the ideal gas law.

ideal gas law constant | The constant the appears in the ideal gas law.

immiscible | Liquids that do not dissolve in each other.

immiscible | Liquids that do not dissolve in each other.

induced-fit model | A model that says an enzyme can undergo a conformational change when it binds substrate molecules.

induced-fit model | A model that says an enzyme can undergo a conformational change when it binds substrate molecules.

inner transition metals | An element in the two rows beneath the main body on the periodic table. Such metals are also called the lanthanide and actinide elements.

inner transition metals | An element in the two rows beneath the main body on the periodic table. Such metals are also called the lanthanide and actinide elements.

inorganic chemistry | The study of the chemistry of all other elements.

inorganic chemistry | The study of the chemistry of all other elements.

intermolecular interactions | A force of attraction between different molecules.

intermolecular interactions | A force of attraction between different molecules.

Ionic bonding | Bonding that results from electrostatic attractions between positively and negatively charged groups.

Ionic bonding | Bonding that results from electrostatic attractions between positively and negatively charged groups.

ionic compounds | A compound formed with an ionic bond.

ionic compounds | A compound formed with an ionic bond.

ionic interactions | An attraction due to ions of opposite charges.

ionic interactions | An attraction due to ions of opposite charges.

irreversible inhibitor | A substance that inactivates an enzyme by bonding covalently to a specific group at the active site.

irreversible inhibitor | A substance that inactivates an enzyme by bonding covalently to a specific group at the active site.

isoelectric point | The pH at which a given amino acid exists in solution as a zwitterion.

isoelectric point | The pH at which a given amino acid exists in solution as a zwitterion.

isomers | Compounds having the same molecular formula but different structural formulas and properties.

isomers | Compounds having the same molecular formula but different structural formulas and properties.

isothermal | A process that occurs at constant temperature.

isothermal | A process that occurs at constant temperature.

isotopes | Atoms of the same element that have different numbers of neutrons.

isotopes | Atoms of the same element that have different numbers of neutrons.

IUPAC System of Nomenclature | A systematic way of naming chemical substances so that each has a unique name.

IUPAC System of Nomenclature | A systematic way of naming chemical substances so that each has a unique name.

joule | The SI unit of energy.

joule | The SI unit of energy.

joule | The SI unit of energy, work, and heat.

joule | The SI unit of energy, work, and heat.

ketogenic amino acids | An amino acid that is converted to acetoacetyl-CoA or acetyl-CoA, which can be used for the synthesis of ketone bodies but not glucose.

ketogenic amino acids | An amino acid that is converted to acetoacetyl-CoA or acetyl-CoA, which can be used for the synthesis of ketone bodies but not glucose.

ketone | An organic compound whose molecules have a carbonyl functional group between two hydrocarbon groups.

ketone | An organic compound whose molecules have a carbonyl functional group between two hydrocarbon groups.

ketoses | A monosaccharide that contains a ketone functional group on the second carbon atom.

ketoses | A monosaccharide that contains a ketone functional group on the second carbon atom.

kinetic theory of gases | The fundamental theory of the behavior of gases.

kinetic theory of gases | The fundamental theory of the behavior of gases.

L sugars | A sugar whose Fischer projection terminates in the same configuration as L-glyceraldehyde.

L sugars | A sugar whose Fischer projection terminates in the same configuration as L-glyceraldehyde.

lattice energy | The strength of interactions between atoms that make ionic bonds.

lattice energy | The strength of interactions between atoms that make ionic bonds.

law | A general statement that explains a large number of observations.

law | A general statement that explains a large number of observations.

law of conservation of matter | In any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant.

law of conservation of matter | In any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant.

Lewis diagrams | A representation that shows valence electrons as dots around the chemical symbol of an atom (also called Lewis electron dot diagrams).

Lewis diagrams | A representation that shows valence electrons as dots around the chemical symbol of an atom (also called Lewis electron dot diagrams).

line-angle formula | An organic chemical formula in which carbon atoms are implied at the corners and ends of lines. Each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds.

line-angle formula | An organic chemical formula in which carbon atoms are implied at the corners and ends of lines. Each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds.

lipids | A compound isolated from body tissues that is more soluble in organic solvents than in water.

lipids | A compound isolated from body tissues that is more soluble in organic solvents than in water.

lock-and-key model | A model that portrays an enzyme as conformationally rigid and able to bond only to a substrate or substrates that exactly fit the active site.

lock-and-key model | A model that portrays an enzyme as conformationally rigid and able to bond only to a substrate or substrates that exactly fit the active site.

mass number | The sum of the numbers of protons and neutrons in a nucleus of an atom.

mass number | The sum of the numbers of protons and neutrons in a nucleus of an atom.

mass-mass calculations | A stoichiometry calculation converting between the mass of one substance and the mass of a different substance in a chemical reaction.

mass-mass calculations | A stoichiometry calculation converting between the mass of one substance and the mass of a different substance in a chemical reaction.

mass/mass percent | A concentration unit that relates the mass of the solute to the mass of the solution.

mass/mass percent | A concentration unit that relates the mass of the solute to the mass of the solution.

mass/volume percent | A concentration unit that relates the mass of the solute to the volume of the solution.

mass/volume percent | A concentration unit that relates the mass of the solute to the volume of the solution.

Matter | Anything that has mass and takes up space.

Matter | Anything that has mass and takes up space.

metabolic pathway | A series of biochemical reactions by which an organism converts a given reactant to a specific end product.

metabolic pathway | A series of biochemical reactions by which an organism converts a given reactant to a specific end product.

microscopic | A view of the universe in which one is working with a few atoms or molecules at a time.

microscopic | A view of the universe in which one is working with a few atoms or molecules at a time.

mitochondria | Small, oval organelles with double membranes; the “power plants” of a cell.

mitochondria | Small, oval organelles with double membranes; the “power plants” of a cell.

modern atomic theory | The fundamental concept that all elements are composed of atoms.

modern atomic theory | The fundamental concept that all elements are composed of atoms.

molar mass | The mass of 1 mol of atoms or molecules.

molar mass | The mass of 1 mol of atoms or molecules.

Molarity | Number of moles of solute per liter of solution.

Molarity | Number of moles of solute per liter of solution.

mole-mass calculations | A stoichiometry calculation converting between masses and moles of different substances in a chemical reaction.

mole-mass calculations | A stoichiometry calculation converting between masses and moles of different substances in a chemical reaction.

mole-mass conversion | The conversion from moles of material to the mass of that same material.

mole-mass conversion | The conversion from moles of material to the mass of that same material.

molecular formulas | A chemical formula for a covalent compound.

molecular formulas | A chemical formula for a covalent compound.

molecular mass | The mass of a molecule, which is the sum of the masses of its atoms.

molecular mass | The mass of a molecule, which is the sum of the masses of its atoms.

molecule | A discrete group of atoms connected by covalent bonds.

molecule | A discrete group of atoms connected by covalent bonds.

mutagens | A chemical or physical agent that cause mutations.

mutagens | A chemical or physical agent that cause mutations.

mutarotation | The ongoing interconversion between anomeric forms of a monosaccharide to form an equilibrium mixture.

mutarotation | The ongoing interconversion between anomeric forms of a monosaccharide to form an equilibrium mixture.

mutation | Any chemical or physical change that alters the nucleotide sequence in DNA.

mutation | Any chemical or physical change that alters the nucleotide sequence in DNA.

neutralization | The reaction of acid and base to make water and a salt.

neutralization | The reaction of acid and base to make water and a salt.

neutron | A subatomic particle with no electric charge.

neutron | A subatomic particle with no electric charge.

nomenclature | The systematic naming of chemical compounds.

nomenclature | The systematic naming of chemical compounds.

nonbonding pairs (or lone pairs) | Electron pair that does not participate in covalent bonds.

nonbonding pairs (or lone pairs) | Electron pair that does not participate in covalent bonds.

noncompetitive inhibitor | A compound that can combine with either the free enzyme or the enzyme-substrate complex at a site distinct from the active site to slow the rate of the reaction.

noncompetitive inhibitor | A compound that can combine with either the free enzyme or the enzyme-substrate complex at a site distinct from the active site to slow the rate of the reaction.

nonelectrolytes | A compound that does not ionize at all when it dissolves.

nonelectrolytes | A compound that does not ionize at all when it dissolves.

nonpolar covalent bond | A covalent bond with a balanced electron distribution across the bond.

nonpolar covalent bond | A covalent bond with a balanced electron distribution across the bond.

Nuclear energy | The controlled harvesting of energy from fission reactions.

Nuclear energy | The controlled harvesting of energy from fission reactions.

nuclear reactor | An apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes.

nuclear reactor | An apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes.

nucleotides | A monomer unit that is linked together to form nucleic acids.

nucleotides | A monomer unit that is linked together to form nucleic acids.

nucleus | The central part of an atom that contains protons and neutrons.

nucleus | The central part of an atom that contains protons and neutrons.

octet rule | The idea that atoms tend to have eight electrons in their valence shell.

octet rule | The idea that atoms tend to have eight electrons in their valence shell.

oil | A triglyceride that is a liquid at room temperature.

oil | A triglyceride that is a liquid at room temperature.

optimum pH | The pH at which a particular enzyme exhibits maximum activity.

optimum pH | The pH at which a particular enzyme exhibits maximum activity.

Organic chemistry | The study of the chemistry of carbon compounds.

Organic chemistry | The study of the chemistry of carbon compounds.

organic compound | A compound containing carbon atoms.

organic compound | A compound containing carbon atoms.

Osmolarity | A way of reporting the total number of particles in a solution to determine the osmotic pressure.

Osmolarity | A way of reporting the total number of particles in a solution to determine the osmotic pressure.

osmotic pressure | The tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal.

osmotic pressure | The tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal.

oxidative deamination | A reaction in which glutamate loses its amino group as an ammonium ion and is oxidized back to α-ketoglutarate.

oxidative deamination | A reaction in which glutamate loses its amino group as an ammonium ion and is oxidized back to α-ketoglutarate.

oxidative phosphorylation | The process that links ATP synthesis to the operation of the electron transport chain.

oxidative phosphorylation | The process that links ATP synthesis to the operation of the electron transport chain.

oxidizing agent | A species that causes oxidation, which is itself reduced.

oxidizing agent | A species that causes oxidation, which is itself reduced.

parts per billion (ppb) | The mass of a solute compared to the mass of a solution times 1,000,000,000.

parts per billion (ppb) | The mass of a solute compared to the mass of a solution times 1,000,000,000.

peptide bond | The amide bond joining two amino acid units in a peptide or protein.

peptide bond | The amide bond joining two amino acid units in a peptide or protein.

period | A row of elements on the periodic table.

period | A row of elements on the periodic table.

periodic table | A chart of elements that groups the elements by some of their properties.

periodic table | A chart of elements that groups the elements by some of their properties.

peripheral proteins | A protein that is more loosely associated with the membrane surface.

peripheral proteins | A protein that is more loosely associated with the membrane surface.

pH scale | A logarithmic scale that relates the concentration of the hydrogen ion in solution.

pH scale | A logarithmic scale that relates the concentration of the hydrogen ion in solution.

phase | A form of matter that has the same physical properties throughout.

phase | A form of matter that has the same physical properties throughout.

phase change | A physical process in which a substance goes from one phase to another.

phase change | A physical process in which a substance goes from one phase to another.

phases | A certain form of matter that includes a specific set of physical properties.

phases | A certain form of matter that includes a specific set of physical properties.

phenols | An aromatic compound with an OH group attached directly to a benzene ring.

phenols | An aromatic compound with an OH group attached directly to a benzene ring.

photosynthesis | The process by which plants use solar energy to convert carbon dioxide and water to glucose.

photosynthesis | The process by which plants use solar energy to convert carbon dioxide and water to glucose.

photosynthesis | The process by which plants use solar energy to convert carbon dioxide and water to glucose.

photosynthesis | The process by which plants use solar energy to convert carbon dioxide and water to glucose.

point mutations | A change in which one nucleotide is substituted, added, or deleted.

point mutations | A change in which one nucleotide is substituted, added, or deleted.

polar | A molecule with a net unequal distribution of electrons in its covalent bonds.

polar | A molecule with a net unequal distribution of electrons in its covalent bonds.

polyamide | A condensation polymer in which the monomer units are joined by an amide linkage.

polyamide | A condensation polymer in which the monomer units are joined by an amide linkage.

polyatomic ions | An ion with more than one atom.

polyatomic ions | An ion with more than one atom.

polycyclic aromatic hydrocarbons (PAHs) | An aromatic hydrocarbon consisting of fused benzene rings sharing a common side.

polycyclic aromatic hydrocarbons (PAHs) | An aromatic hydrocarbon consisting of fused benzene rings sharing a common side.

polymers | A giant molecule formed by the combination of monomers in a repeating manner.

polymers | A giant molecule formed by the combination of monomers in a repeating manner.

polypeptides | A chain of about 50 or more amino acids.

polypeptides | A chain of about 50 or more amino acids.

polysaccharides | A carbohydrate containing many monosaccharide units.

polysaccharides | A carbohydrate containing many monosaccharide units.

polyunsaturated fatty acids | A fatty acid that has two or more carbon-to-carbon double bonds.

polyunsaturated fatty acids | A fatty acid that has two or more carbon-to-carbon double bonds.

power | The exponent in a number expressed in scientific notation.

power | The exponent in a number expressed in scientific notation.

pressure | Force divided by area.

pressure | Force divided by area.

primary (1°) alcohol | A compound with an OH group on a carbonatom that is attached to only one other carbon atom.

primary (1°) alcohol | A compound with an OH group on a carbonatom that is attached to only one other carbon atom.

primary structure | The sequence of amino acids in a polypeptide chain or protein.

primary structure | The sequence of amino acids in a polypeptide chain or protein.

products | A substance on the right side of the arrow in a chemical equation.

products | A substance on the right side of the arrow in a chemical equation.

Proteins | A compound of high molar mass consisting largely or entirely of amino acids linked together.

Proteins | A compound of high molar mass consisting largely or entirely of amino acids linked together.

proton | A subatomic particle with a positive charge.

proton | A subatomic particle with a positive charge.

purines | A heterocyclic amine consisting of a pyrimidine ring fused to a five-member ring with two nitrogen atoms.

purines | A heterocyclic amine consisting of a pyrimidine ring fused to a five-member ring with two nitrogen atoms.

quantized | Having a fixed value.

quantized | Having a fixed value.

quantum mechanics | The modern theory of electron behavior.

quantum mechanics | The modern theory of electron behavior.

quaternary structure | The arrangement of multiple subunits in a protein.

quaternary structure | The arrangement of multiple subunits in a protein.

rad | A unit of radioactive exposure equal to 0.01 J/g of tissue.

rad | A unit of radioactive exposure equal to 0.01 J/g of tissue.

radioactivity | Emanations of particles and radiation from atomic nuclei.

radioactivity | Emanations of particles and radiation from atomic nuclei.

reducing sugar | Any carbohydrate capable of reducing a mild oxidizing agent, such as Tollens' or Benedict's reagents, without first undergoing hydrolysis.

reducing sugar | Any carbohydrate capable of reducing a mild oxidizing agent, such as Tollens' or Benedict's reagents, without first undergoing hydrolysis.

rem | A unit of radioactive exposure that includes a factor to account for the type of radioactivity.

rem | A unit of radioactive exposure that includes a factor to account for the type of radioactivity.

replication | The process in which the DNA in a dividing cell is copied.

replication | The process in which the DNA in a dividing cell is copied.

respiration | The biochemical process by which the oxygen we inhale oxidizes foodstuffs to carbon dioxide and water.

respiration | The biochemical process by which the oxygen we inhale oxidizes foodstuffs to carbon dioxide and water.

Respiration | The process by which cells oxidize organic molecules in the presence of gaseous oxygen to produce carbon dioxide, water, and energy in the form of ATP.

Respiration | The process by which cells oxidize organic molecules in the presence of gaseous oxygen to produce carbon dioxide, water, and energy in the form of ATP.

retroviruses | An RNA virus that directs the synthesis of a DNA copy in the host cell.

retroviruses | An RNA virus that directs the synthesis of a DNA copy in the host cell.

Ribonucleic acid (RNA) | The nucleic acid responsible for using the genetic information encoded in DNA.

Ribonucleic acid (RNA) | The nucleic acid responsible for using the genetic information encoded in DNA.

Ribosomes | A cellular substructure where proteins are synthesized.

Ribosomes | A cellular substructure where proteins are synthesized.

round | The process of assessing the final significant figure of a quantity to determine if it should be kept or moved higher.

round | The process of assessing the final significant figure of a quantity to determine if it should be kept or moved higher.

saponification | The hydrolysis of fats and oils in the presence of a base to make soap.

saponification | The hydrolysis of fats and oils in the presence of a base to make soap.

saponification | The hydrolysis of fats and oils in the presence of a base to make soap.

saponification | The hydrolysis of fats and oils in the presence of a base to make soap.

Science | The process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations.

Science | The process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations.

scientific method | An organized procedure for learning answers to questions.

scientific method | An organized procedure for learning answers to questions.

secondary (2°) alcohol | A compound with an OH group on a carbon atom that is attached to two other carbon atoms.

secondary (2°) alcohol | A compound with an OH group on a carbon atom that is attached to two other carbon atoms.

secondary structure | The fixed arrangement of the polypeptide backbone.

secondary structure | The fixed arrangement of the polypeptide backbone.

semimetals (or metalloids) | An element whose properties are intermediate between metals and nonmetals.

semimetals (or metalloids) | An element whose properties are intermediate between metals and nonmetals.

shells | A grouping of electrons within an atom.

shells | A grouping of electrons within an atom.

significant figures | All the digits of a measured quantity known with certainty and the first uncertain, or estimated, digit.

significant figures | All the digits of a measured quantity known with certainty and the first uncertain, or estimated, digit.

single bond | A covalent bond formed by a single pair of electrons.

single bond | A covalent bond formed by a single pair of electrons.

solubility | The limit of how much solute can be dissolved in a given amount of solvent.

solubility | The limit of how much solute can be dissolved in a given amount of solvent.

solute | The minor component of a solution.

solute | The minor component of a solution.

solution | Another name for a homogeneous mixture.

solution | Another name for a homogeneous mixture.

specific heat | A proportionality constant that relates heat to a temperature change.

specific heat | A proportionality constant that relates heat to a temperature change.

Sphingolipids | A lipid that contains the unsaturated amino alcohol sphingosine.

Sphingolipids | A lipid that contains the unsaturated amino alcohol sphingosine.

Sphingomyelins | A sphingolipid that contains a fatty acid unit, a phosphoric acid unit, a sphingosine unit, and a choline unit.

Sphingomyelins | A sphingolipid that contains a fatty acid unit, a phosphoric acid unit, a sphingosine unit, and a choline unit.

spontaneous fission (or fission) | The breaking apart of an atomic nucleus into smaller nuclei.

spontaneous fission (or fission) | The breaking apart of an atomic nucleus into smaller nuclei.

standard temperature and pressure (STP) | 273 K (0°C) and 1.00 atm pressure.

standard temperature and pressure (STP) | 273 K (0°C) and 1.00 atm pressure.

steroids | A lipid with a four-fused-ring structure.

steroids | A lipid with a four-fused-ring structure.

Stock system | The system of indicating a cation's charge with roman numerals.

Stock system | The system of indicating a cation's charge with roman numerals.

stoichiometry | The study of the numerical relationships between the reactants and the products in a balanced chemical equation.

stoichiometry | The study of the numerical relationships between the reactants and the products in a balanced chemical equation.

strong acid | An acid that is 100% ionized in aqueous solution.

strong acid | An acid that is 100% ionized in aqueous solution.

strong base | A base that is 100% ionized in aqueous solution.

strong base | A base that is 100% ionized in aqueous solution.

structural formula | A chemical formula that shows how the atoms of a molecule are attached to one another.

structural formula | A chemical formula that shows how the atoms of a molecule are attached to one another.

subshells | A grouping of electrons within a shell.

subshells | A grouping of electrons within a shell.

substrate-level phosphorylation | The synthesis of ATP by the direct transfer of a phosphate group from a metabolite to ADP.

substrate-level phosphorylation | The synthesis of ATP by the direct transfer of a phosphate group from a metabolite to ADP.

substrates | A compound on which an enzyme acts.

substrates | A compound on which an enzyme acts.

tertiary (3°) alcohol | A compond with an OH group on a carbon atom that is attached to three other carbon atoms.

tertiary (3°) alcohol | A compond with an OH group on a carbon atom that is attached to three other carbon atoms.

tertiary (3°) amine | A compound that has three alkyl or aryl groups on the nitrogen atom.

tertiary (3°) amine | A compound that has three alkyl or aryl groups on the nitrogen atom.

Tertiary structure | The unique three-dimensional shape of a polypeptide chain as a whole.

Tertiary structure | The unique three-dimensional shape of a polypeptide chain as a whole.

theory | A general statement that describes a large set of observations and data.

theory | A general statement that describes a large set of observations and data.

Thiols | A compound with an SH functional group.

Thiols | A compound with an SH functional group.

torr | Another name for millimeters of mercury.

torr | Another name for millimeters of mercury.

tracer | A substance that can be used to follow the pathway of that substance through some structure.

tracer | A substance that can be used to follow the pathway of that substance through some structure.

Transamination | An exchange of functional groups between any amino acid and an α-keto acid.

Transamination | An exchange of functional groups between any amino acid and an α-keto acid.

transcription | The process in which RNA is synthesized from a DNA template.

transcription | The process in which RNA is synthesized from a DNA template.

translation | The process in which the information encoded in mRNA is used to direct the sequencing of amino acids to synthesize a protein.

translation | The process in which the information encoded in mRNA is used to direct the sequencing of amino acids to synthesize a protein.

triglycerides | An ester composed of three fatty acid units linked to glycerol and found in fats and oils.

triglycerides | An ester composed of three fatty acid units linked to glycerol and found in fats and oils.

triple bonds | Three pairs of electrons being shared by two atoms in a molecule.

triple bonds | Three pairs of electrons being shared by two atoms in a molecule.

unit | The scale of measurement for a quantity.

unit | The scale of measurement for a quantity.

unsaturated | A solution whose solute is less than its solubility limit.

unsaturated | A solution whose solute is less than its solubility limit.

unsaturated hydrocarbons | An alkene or alkyne having one or more multiple (double or triple) bonds between carbon atoms.

unsaturated hydrocarbons | An alkene or alkyne having one or more multiple (double or triple) bonds between carbon atoms.

valence shell electron pair repulsion (VSEPR) | The general concept that estimates the shape of a simple molecule.

valence shell electron pair repulsion (VSEPR) | The general concept that estimates the shape of a simple molecule.

vapor pressure | The pressure of a vapor that is in equilibrium with its liquid phase.

vapor pressure | The pressure of a vapor that is in equilibrium with its liquid phase.

vapor pressure depression | The lowering of the vapor pressure of a solution versus the pure solvent.

vapor pressure depression | The lowering of the vapor pressure of a solution versus the pure solvent.

Viruses | An infectious agent that is much smaller and simpler than bacteria.

Viruses | An infectious agent that is much smaller and simpler than bacteria.

Vitamins | An organic compound that is essential in very small amounts for the maintenance of normal metabolism.

Vitamins | An organic compound that is essential in very small amounts for the maintenance of normal metabolism.

Volume | The amount of space that a given substance occupies.

Volume | The amount of space that a given substance occupies.

volume/volume percent | A concentration unit that relates the volume of the solute to the volume of the solution.

volume/volume percent | A concentration unit that relates the volume of the solute to the volume of the solution.

weak base | A base that is less than 100% ionized in aqueous solution.

weak base | A base that is less than 100% ionized in aqueous solution.

zwitterion | An electrically neutral compound that contains both negatively and positively charged groups.

zwitterion | An electrically neutral compound that contains both negatively and positively charged groups.

β-oxidation | A sequence of four reactions in which fatty acyl-CoA molecules are oxidized, leading to the removal of acetyl-CoA molecules.

β-oxidation | A sequence of four reactions in which fatty acyl-CoA molecules are oxidized, leading to the removal of acetyl-CoA molecules.

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