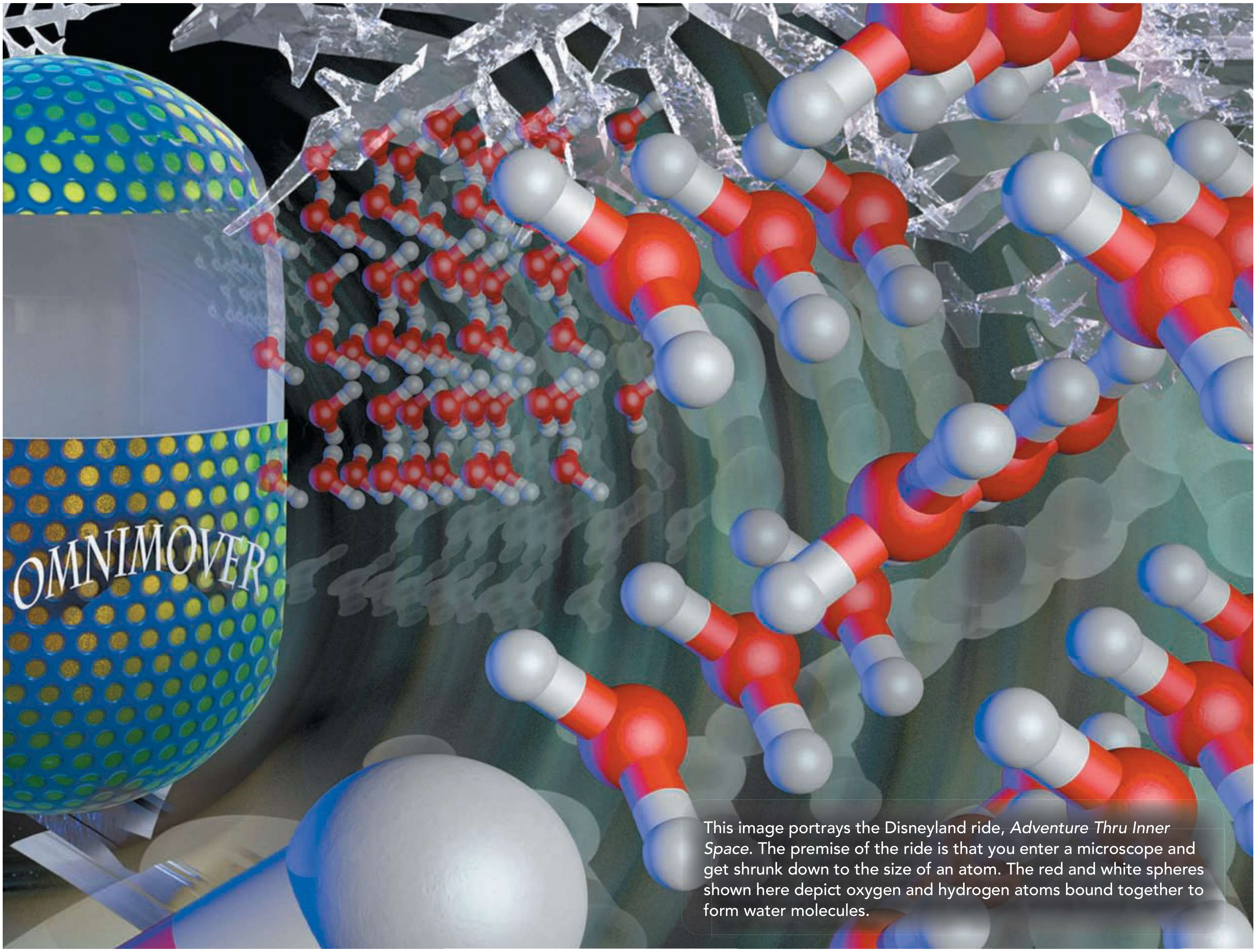


This image portrays the Disneyland ride, *Adventure Thru Inner Space*. The premise of the ride is that you enter a microscope and get shrunk down to the size of an atom. The red and white spheres shown here depict oxygen and hydrogen atoms bound together to form water molecules.

1.1 Atoms and Molecules

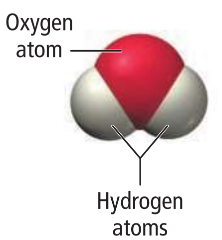
WATCH **NOW!**

KEY CONCEPT VIDEO 1.1
Atoms and Molecules

As I sat in the "omnimover" and listened to the narrator's voice telling me that I was shrinking down to the size of an atom, I grew apprehensive but curious. Just minutes before, while waiting in line, I witnessed what appeared to be full-sized humans entering a microscope and emerging from the other end many times smaller. I was seven years old, and I was about to ride *Adventure Thru Inner Space*, a Disneyland ride (in Tomorrowland) that simulated what it would be like to shrink to the size of an atom. The ride began with darkness and shaking, but then the shaking stopped and giant snowflakes appeared. The narrator explained that you were in the process of shrinking to an ever-smaller size (which explains why the snowflakes grew larger and larger). Soon, you entered the wall of the snowflake itself and began to see water molecules all around you. These also grew larger as you continued your journey into inner space and eventually ended up within the atom itself. Although this Disneyland ride bordered on being corny, and although it has since been shut down, it was my favorite ride as a young child.

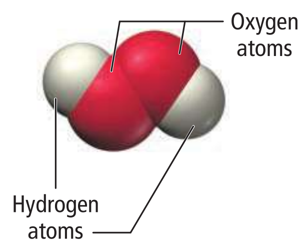
That ride sparked my interest in the world of atoms and molecules, an interest that has continued and grown to this day. I am a chemist because I am obsessed with the connection between the "stuff" around us and the atoms and molecules that compose that stuff. More specifically, I love the idea that we humans have been able to figure out the connection between the *properties of the stuff* around us and the *properties of atoms and molecules*. **Atoms** are submicroscopic particles that are the fundamental building blocks of ordinary matter. Free atoms are rare in nature; instead they bind together in specific geometrical arrangements to form **molecules**. A good example of a molecule is the water molecule, which I remember so well from the Disneyland ride.

Water molecule



A water molecule is composed of one oxygen atom bound to two hydrogen atoms in the shape shown at left. The exact properties of the water molecule—the atoms that compose it, the distances between those atoms, and the geometry of how the atoms are bound together—determine the properties of water. If the molecule were different, water would be different. For example, if water contained two oxygen atoms instead of just one, it would be a molecule like this:

Hydrogen peroxide molecule



The hydrogen peroxide we use as an antiseptic or bleaching agent is considerably diluted.

This molecule is hydrogen peroxide, which you may have encountered if you have ever bleached your hair. A hydrogen peroxide molecule is composed of *two* oxygen atoms and two hydrogen atoms. This seemingly small molecular difference results in a huge difference in the properties of water and hydrogen peroxide. Water is the familiar and stable liquid we all drink and bathe in. Hydrogen peroxide, in contrast, is an unstable liquid that, in its pure form, burns the skin on contact and is used in rocket fuel. When you pour water onto your hair, your hair simply becomes wet. However, if you put diluted hydrogen peroxide on your hair, a chemical reaction occurs that strips your hair of its color.

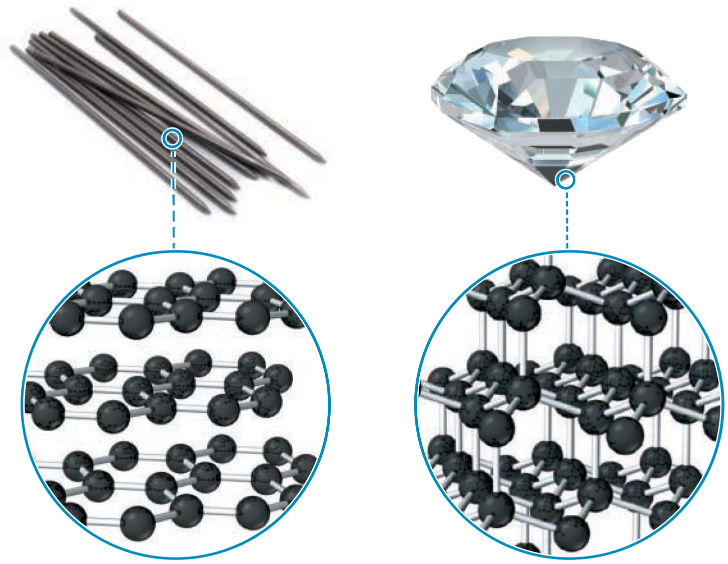
The details of how specific atoms bond to form a molecule—in a straight line, at a particular angle, in a ring, or in some other pattern—as well as the type of atoms in the molecule, determine everything about the substance that the molecule composes. If we want to understand the substances around us, we must understand the atoms and molecules that compose them—this is the central goal of chemistry. A good simple definition of **chemistry** is

> **Chemistry—the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules.**

The term *atoms* in this definition can be interpreted loosely to include atoms that have lost or gained electrons.

Throughout this book, we explore the connection between atoms and molecules and the matter they compose. We seek to understand how differences on the atomic or molecular level affect the properties on the macroscopic level. Before we move on, let's examine one more example that demonstrates this principle. Consider the structures of graphite and diamond.

Graphite is the slippery black substance (often called pencil lead) that you have probably used in a mechanical pencil. Diamond is the brilliant gemstone found in jewelry. Graphite and diamond are both composed of exactly the same atoms—carbon atoms. The striking differences between the substances are a result of how those atoms are arranged. In graphite, the atoms are arranged in sheets. The atoms within each sheet are tightly bound to each other, but the sheets are *not* tightly bound to other sheets. Therefore the sheets can slide past each other, which is why the graphite in a pencil leaves a trail as you write. In diamond, by contrast, the carbon atoms are all bound together in a three-dimensional structure where layers are strongly bound to other layers, resulting in the strong, nearly unbreakable substance. This example illustrates how even the same atoms can compose vastly different substances when they are bound together in different patterns. Such is the atomic and molecular world—small differences in atoms and molecules can result in large differences in the substances that they compose.



Graphite structure Diamond structure

1.2 The Scientific Approach to Knowledge

Throughout history, humans have approached knowledge about the physical world in different ways. For example, the Greek philosopher Plato (427–347 B.C.E.) thought that the best way to learn about reality was—not through the senses—but through reason. He believed that the physical world was an imperfect representation of a perfect and transcendent world (a world beyond space and time). For him, true knowledge came, not through observing the real physical world, but through reasoning and thinking about the ideal one.

Although some Greek philosophers, such as Aristotle, did use observation to attain knowledge, they did not emphasize experiment and measurement to the extent that modern science does.

The *scientific* approach to knowledge, however, is exactly the opposite of Plato's. Scientific knowledge is empirical—it is based on *observation* and *experiment*. Scientists observe and perform experiments on the physical world to learn about it. Some observations and experiments are qualitative (noting or describing how a process happens), but many are quantitative (measuring or quantifying something about the process). For example, Antoine Lavoisier (1743–1794), a French chemist who studied combustion (burning), made careful measurements of the mass of objects before and after burning them in closed containers. He noticed that there was no change in the total mass of material within the container during combustion. In doing so, Lavoisier made an important *observation* about the physical world.

▲ French chemist Antoine Lavoisier with his wife, Marie, who helped him in his work by illustrating his experiments and translating scientific articles from English. Lavoisier, who also made significant contributions to agriculture, industry, education, and government administration, was executed during the French Revolution.
(The Metropolitan Museum of Art)

Observations often lead scientists to formulate a **hypothesis**, a tentative interpretation or explanation of the observations. For example, Lavoisier explained his observations on combustion by hypothesizing that when a substance burns, it combines with a component of air. A good hypothesis is *falsifiable*, which means that it makes predictions that can be confirmed or refuted by further observations. Scientists test hypotheses by **experiments**, highly controlled procedures designed to generate observations that confirm or refute a hypothesis. The results of an experiment may support a hypothesis or prove it wrong—in which case the scientist must modify or discard the hypothesis.

In some cases, a series of similar observations leads to the development of a **scientific law**, a brief statement that summarizes past observations and predicts future ones. Lavoisier summarized his observations on combustion with the **law of conservation of mass**, which states, "In a chemical reaction, matter is neither created nor destroyed." This statement summarized his observations on chemical reactions and predicted the outcome of future observations on reactions. Laws, like hypotheses, are also subject to experiments, which can support them or prove them wrong.

Scientific laws are not *laws* in the same sense as civil or governmental laws. Nature does not follow laws in the way that we obey the laws against speeding or running a stop sign. Rather, scientific laws *describe* how nature behaves—they are generalizations about what nature does. For that reason, some people find it more appropriate to refer to them as *principles* rather than *laws*.

One or more well-established hypotheses may form the basis for a scientific **theory**. A scientific theory is a model for the way nature is and tries to explain not merely what nature does but why. As such, well-established theories are the pinnacle of scientific knowledge, often predicting behavior far beyond the observations or laws from which they were developed. A good example of a theory is the **atomic theory** proposed by English chemist John Dalton (1766–1844). Dalton explained the law of conservation of mass, as well as other laws and observations of the time, by proposing that matter is composed of small, indestructible particles called atoms. Since these particles are merely rearranged in chemical changes (and not created or destroyed), the total amount of mass remains the same. Dalton's theory is a model for the physical world—it gives us insight into how nature works and, therefore, *explains* our laws and observations.

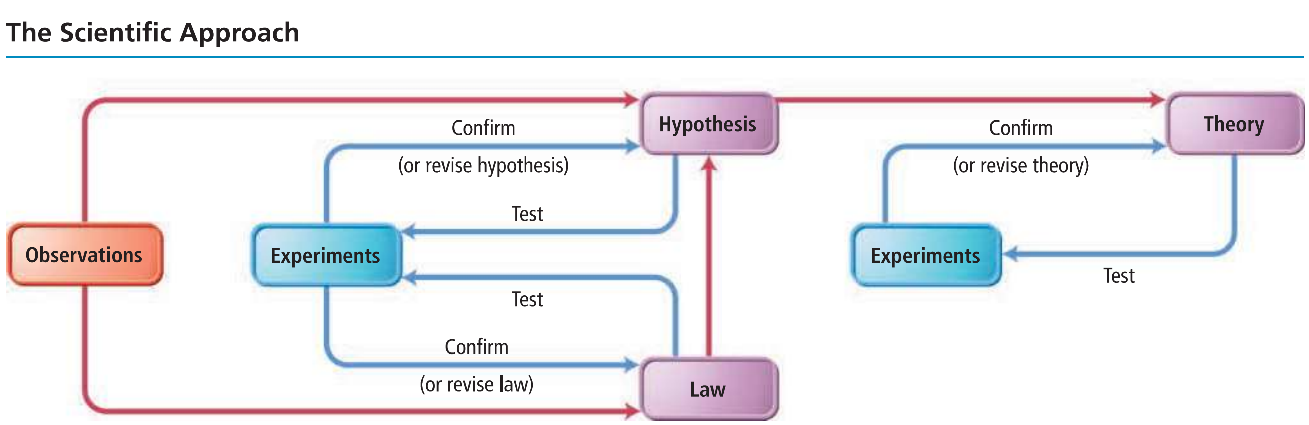
In Dalton's time, people thought atoms were indestructible. Today, because of nuclear reactions, we know that atoms can be broken apart into their smaller components.

Finally, the scientific approach returns to observation to test theories. For example, scientists can test the atomic theory by trying to isolate single atoms or by trying to image them (both of which, by the way, have already been accomplished). Theories are validated by experiments; however, theories can never be conclusively proven because some new observation or experiment always has the potential to reveal a flaw. Notice that the scientific approach to knowledge begins with observation and ends with observation. An experiment is in essence a highly controlled procedure for generating critical observations designed to test a theory or hypothesis. Each new set of observations has the potential to refine the original model. Figure 1.1▼ summarizes one way to map the scientific approach to knowledge. Scientific laws, hypotheses, and theories are all subject to continued experimentation. If a law, hypothesis, or theory is proved wrong by an experiment, it must be revised and tested with new experiments. Over time, the scientific community eliminates or corrects poor theories and laws, and valid theories and laws—those consistent with experimental results—remain.

Established theories with strong experimental support are the most powerful pieces of scientific knowledge. You may have heard the phrase "That is just a theory," as if theories are easily dismissible. Such a statement reveals a deep misunderstanding of the nature of a scientific theory. Well-established theories are as close to truth as we get in science. The idea that all matter is made of atoms is "just a theory," but it has over 200 years of experimental evidence to support it. It is a powerful piece of scientific knowledge on which many other scientific ideas are based.

One last word about the scientific approach to knowledge: some people wrongly imagine science to be a strict set of rules and procedures that automatically leads to inarguable, objective facts. This is not the case. Even our diagram of the scientific approach to knowledge is only an idealization of real science, useful to help us see the key distinctions of science. Real science requires hard work, care, creativity, and even a bit of luck.

▼ FIGURE 1.1 The Scientific Approach to Knowledge



Scientific theories do not just arise out of data—men and women of genius and creativity craft theories. A great theory is not unlike a master painting, and many see a similar kind of beauty in both. (For more on this aspect of science, see the accompanying box entitled *Thomas S. Kuhn and Scientific Revolutions*.)

LAWS AND THEORIES Which statement best explains the difference between a law and a theory?

(a) A law is truth; a theory is mere speculation.
(b) A law summarizes a series of related observations; a theory gives the underlying reasons for them.
(c) A theory describes *what* nature does; a law describes *why* nature does it.

1.1 Cc Conceptual Connection

ANSWER **NOW!**

THE NATURE OF SCIENCE | Thomas S. Kuhn and Scientific Revolutions

When scientists talk about science, they often talk in ways that imply that theories are "true." Further, they talk as if they arrive at theories in logical and unbiased ways. For example, a theory central to chemistry that we have discussed in this chapter is John Dalton's atomic theory—the idea that all matter is composed of atoms. Is this theory "true"? Was it reached in logical, unbiased ways? Will this theory still be around in 200 years?

The answers to these questions depend on how we view science and its development. One way to view science—let's call it the *traditional view*—is as the continual accumulation of knowledge and the building of increasingly precise theories. In this view, a scientific theory is a model of the world that reflects what is *actually in* nature. New observations and experiments result in gradual adjustments to theories. Over time, theories get better, giving us a more accurate picture of the physical world.

In the twentieth century, a different view of scientific knowledge began to develop. A book by Thomas Kuhn (1922–1996), published in 1962 and entitled *The Structure of Scientific Revolutions*, challenged the traditional view. Kuhn's ideas came from his study of the history of science, which, he argued, does not support the idea that science progresses in a smooth, cumulative way. According to Kuhn, science goes through fairly quiet periods that he called *normal science*. In these periods, scientists make their data fit the reigning theory, or paradigm. Small inconsistencies are swept aside during periods of normal science. However, when too many inconsistencies and anomalies develop, a crisis emerges. The crisis brings about a *revolution* and a new reigning theory. According to Kuhn, the new theory is usually quite different from the old one; it not only helps us to make sense of new or anomalous information, but it also enables us to see accumulated data from the past in a dramatically new way.

Kuhn further contended that theories are held for reasons that are not always logical or unbiased, and that theories are not *true* models—in the sense of a one-to-one mapping—of the physical world. Because new theories are often so different from the ones they replace, he argued, and because old theories always make good sense to those holding them, they must not be "True" with a capital *T*; otherwise "truth" would be constantly changing.

Kuhn's ideas created a controversy among scientists and science historians that continues to this day. Some, especially postmodern philosophers of science, have taken Kuhn's ideas one step further. They argue that scientific knowledge is *completely* biased and lacks any objectivity. Most scientists, including Kuhn, would disagree. Although Kuhn pointed out that scientific knowledge has *arbitrary elements*, he also said, "*Observation . . . can and must drastically restrict the range of admissible scientific belief, else there would be no science.*" In other words, saying that science contains arbitrary elements is quite different from saying that science itself is arbitrary.

QUESTION *In his book, Kuhn stated, "A new theory . . . is seldom or never just an increment to what is already known." From your knowledge of the history of science, can you think of any examples that support Kuhn's statement? Do you know of any instances in which a new theory or model was drastically different from the one it replaced?*

1.3 The Classification of Matter

WATCH **NOW!**
KEY CONCEPT VIDEO 1.3
Classifying Matter

Matter is anything that occupies space and has mass. Your desk, your chair, and even your body are all composed of matter. Less obviously, the air around you is also matter—it too occupies space and has mass. We call a specific instance of matter—such as air, water, or sand—a **substance**. We classify matter according to its **state** (its physical form) and its **composition** (the basic components that make it up).

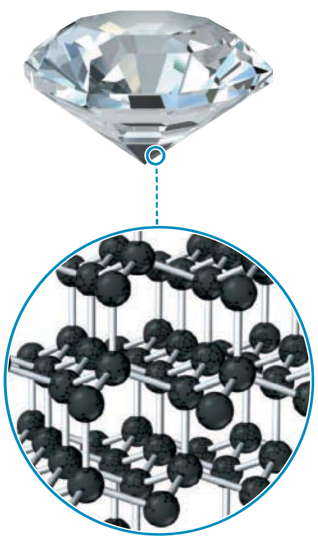
The States of Matter: Solid, Liquid, and Gas

Matter exists in three different states: **solid**, **liquid**, and **gas**. In *solid matter*, atoms or molecules pack closely to each other in fixed locations. Although the atoms and molecules in a solid vibrate, they do not move around or past each other. Consequently, a solid has a fixed volume and rigid shape. Ice, aluminum, and diamond are examples of solids. Solid matter may be **crystalline**, in which case its atoms or molecules are in patterns with long-range, repeating order (Figure 1.2▼), or it may be **amorphous**, in which case its atoms or molecules do not have any long-range order. Table salt and diamond are examples of *crystalline* solids; the well-ordered geometric shapes of salt and diamond crystals reflect the well-ordered geometric arrangement of their atoms (although this is not the case for *all* crystalline solids). Examples of *amorphous* solids include glass and plastic. In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other, giving liquids a fixed volume but not a fixed shape. Liquids assume the shape of their containers. Water, alcohol, and gasoline are all substances that are liquids at room temperature.

The state of matter changes from solid to liquid to gas with increasing temperature.

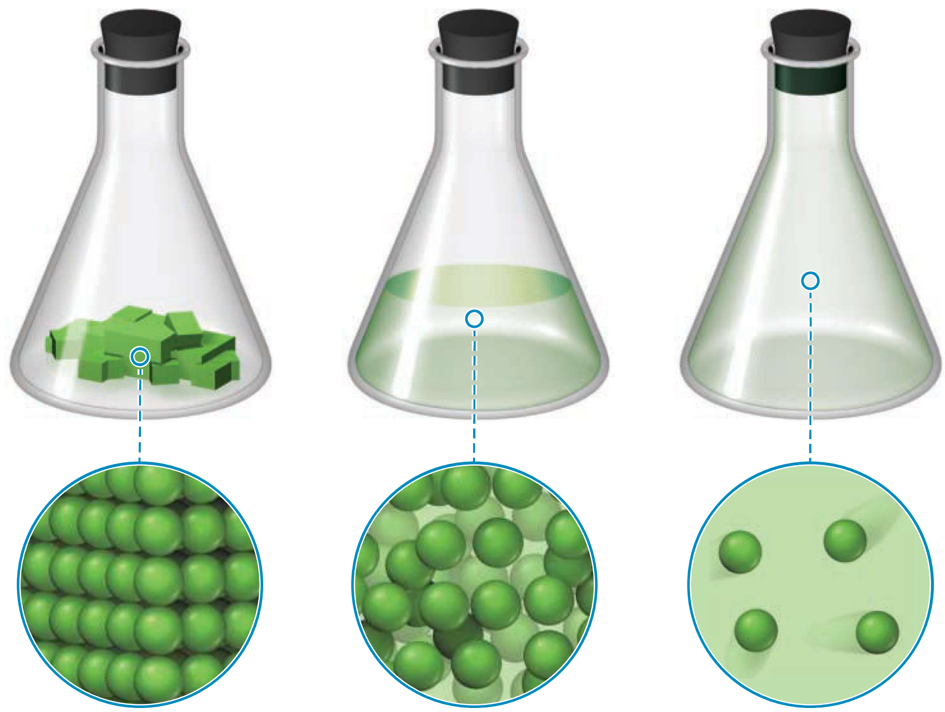
Glass and other amorphous solids can be thought of, from one point of view, as intermediate between solids and liquids. Their atoms are fixed in position at room temperature, but they have no long-range structure and do not have distinct melting points.

Crystalline Solid:
Atoms are arranged in a regular three-dimensional pattern



Diamond
C (*s*, diamond)

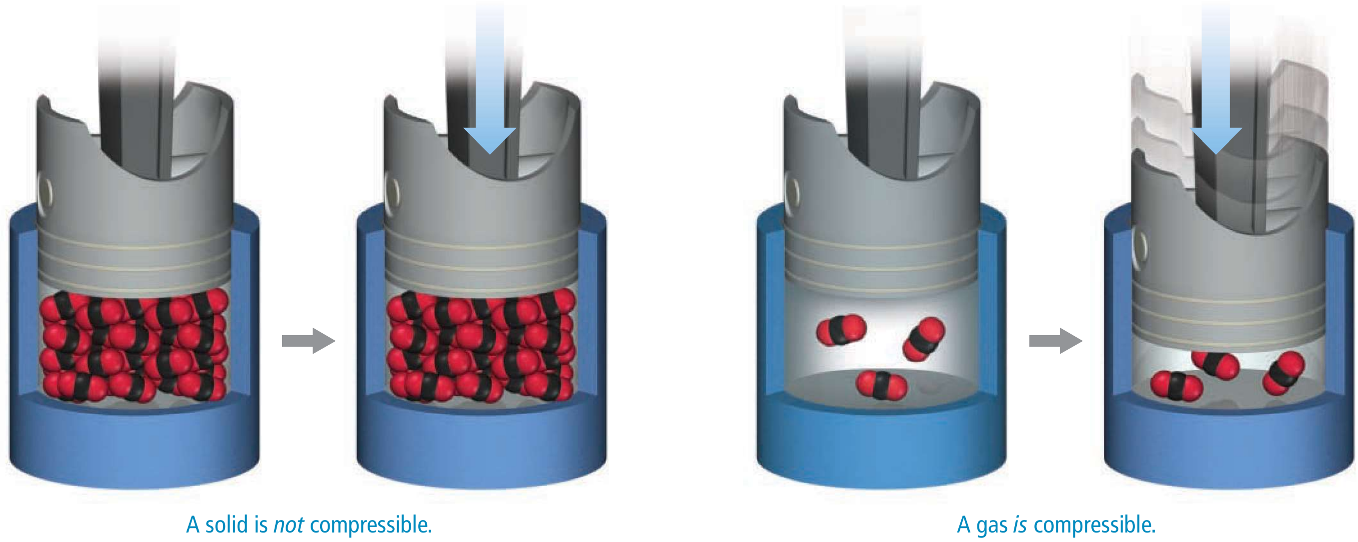
▲ **FIGURE 1.2 Crystalline Solid** Diamond (first discussed in Section 1.1) is a crystalline solid composed of carbon atoms arranged in a regular, repeating pattern.



Solid matter **Liquid matter** **Gaseous matter**

▲ In a solid, the atoms or molecules are fixed in place and can only vibrate. In a liquid, although the atoms or molecules are closely packed, they can move past one another, allowing the liquid to flow and assume the shape of its container. In a gas, the atoms or molecules are widely spaced, making gases compressible as well as fluid (able to flow).

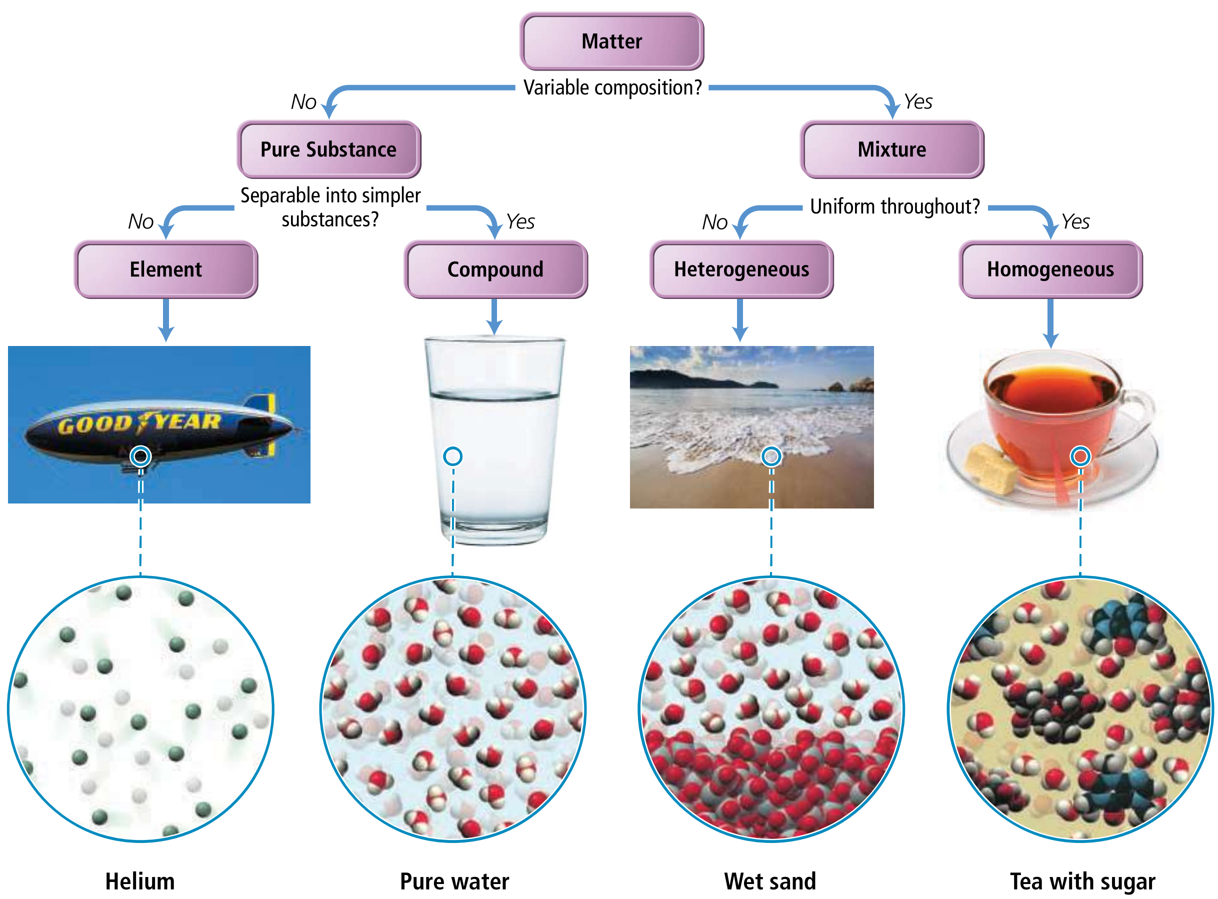
In *gaseous matter*, atoms or molecules have a lot of space between them and are free to move relative to one another, making gases *compressible* (Figure 1.3▶). When you squeeze a balloon or sit down on an air mattress, you force the atoms and molecules into a smaller space so that they are closer together. Gases always assume the shape *and* volume of their containers. Substances that are gases at room temperature include helium, nitrogen (the main component of air), and carbon dioxide.



▲ **FIGURE 1.3 The Compressibility of Gases** Gases can be compressed—squeezed into a smaller volume—because there is so much empty space between atoms or molecules in the gaseous state.

Classifying Matter by Composition: Elements, Compounds, and Mixtures

In addition to classifying matter according to its state, we classify it according to its composition, as shown in the following chart:



The first division in the classification of matter is between a *pure substance* and a *mixture*. A **pure substance** is made up of only one component, and its composition is invariant (it does not vary from one sample to another). The *components* of a pure substance can be individual atoms or groups of atoms joined together. For example, helium, water, and table salt (sodium chloride) are all pure substances. Each of these substances is made up of only one component: helium is made up of helium atoms, water is made up of water molecules, and sodium chloride is made up of sodium chloride units. The composition of a pure sample of any one of these substances is always exactly the same (because you can't vary the composition of a substance made up of only one component).

A **mixture**, by contrast, is composed of two or more components in proportions that can vary from one sample to another. For example, sweetened tea, composed primarily of water molecules and sugar molecules (with a few other substances mixed in), is a mixture. We can make tea slightly sweet (a small proportion of sugar to water) or very sweet (a large proportion of sugar to water) or any level of sweetness in between.

We categorize pure substances themselves into two types—*elements* and *compounds*—depending on whether or not they can be broken down (or decomposed) into simpler substances. Helium, which we just noted is a pure substance, is also a good example of an **element**, a substance that cannot be chemically broken down into simpler substances. Water, also a pure substance, is a good example of a **compound**, a substance composed of two or more elements (in this case, hydrogen and oxygen) in a fixed, definite proportion. On Earth, compounds are more common than pure elements because most elements combine with other elements to form compounds.

All known elements are listed in the periodic table in the inside front cover of this book.

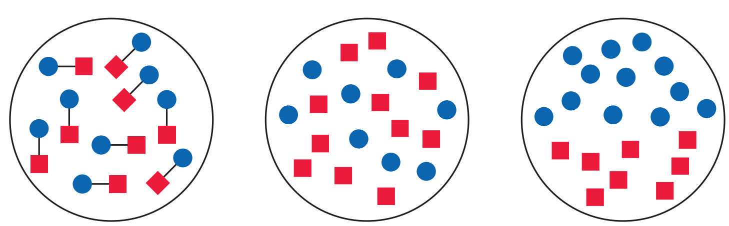
We also categorize mixtures into two types—heterogeneous and homogeneous—depending on how *uniformly* the substances within them mix. Wet sand is a **heterogeneous mixture**, one in which the composition varies from one region of the mixture to another. Sweetened tea is a **homogeneous mixture**, one with the same composition throughout. Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly. Heterogeneous mixtures are made up of distinct regions because the atoms or molecules that compose them separate. Here again we see that the properties of matter are determined by the atoms or molecules that compose it.

Classifying a substance according to its composition is not always obvious and requires that we either know the true composition of the substance or are able to test it in a laboratory. For now, we focus on relatively common substances that you are likely to have encountered. Throughout this course, you will gain the knowledge to understand the composition of a larger variety of substances.

ANSWER **NOW!**

1.2 Cc Conceptual Connection

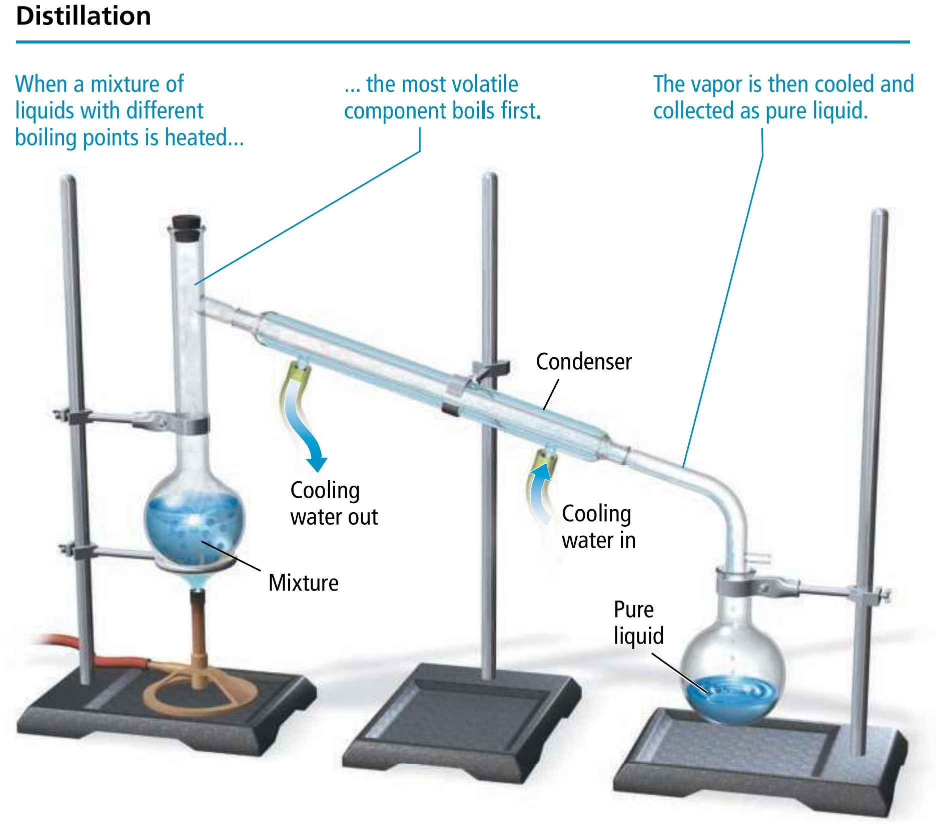
PURE SUBSTANCES AND MIXTURES In these images, a blue circle represents an atom of one type of element, and a red square represents an atom of a second type of element. Which image is a pure substance?



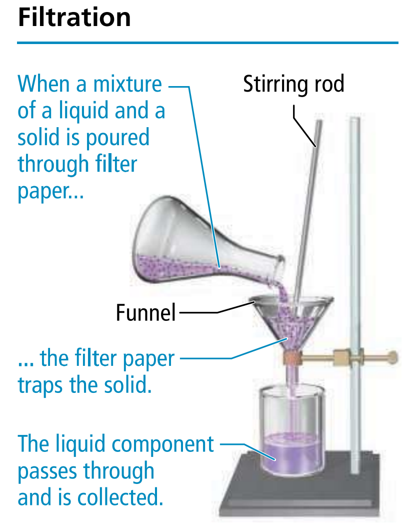
(a) **(b)** **(c)** None of the these **(d)**

Separating Mixtures

Chemists often want to separate a mixture into its components. Such separations can be easy or difficult, depending on the components in the mixture. In general, mixtures are separable because the different components have different physical or chemical properties. We can use various techniques that exploit these differences to achieve

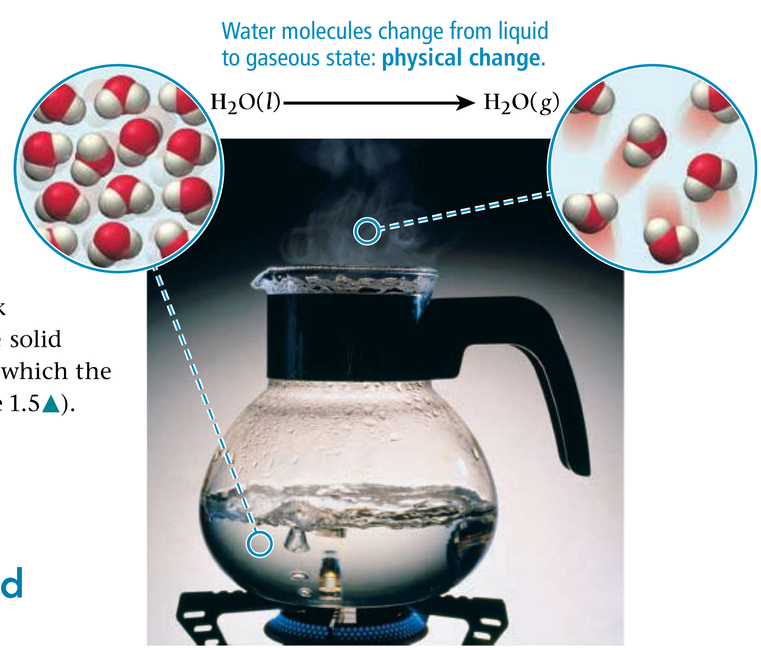


▲FIGURE 1.4 **Separating Substances by Distillation**



▲ FIGURE 1.5 **Separating Substances by Filtration**

separation. For example, we can separate a mixture of sand and water by **decanting**—carefully pouring off—the water into another container. A homogeneous mixture of liquids can usually be separated by **distillation**, a process in which the mixture is heated to boil off the more **volatile** (easily vaporizable) liquid. The volatile liquid is then recondensed in a condenser and collected in a separate flask (Figure 1.4▲). If a mixture is composed of an insoluble solid and a liquid, we can separate the two by **filtration**, in which the mixture is poured through filter paper in a funnel (Figure 1.5▲).



▲ FIGURE 1.6 **Boiling, a Physical Change** When water boils, it turns into a gas but does not alter its chemical identity—the water molecules are the same in both the liquid and gaseous states. Boiling is a physical change, and the boiling point of water is a physical property.

1.4 Physical and Chemical Changes and Physical and Chemical Properties

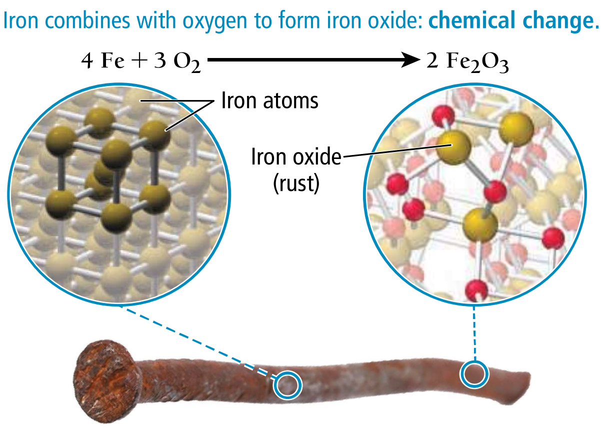
Every day we witness changes in matter: ice melts, iron rusts, gasoline burns, fruit ripens, and water evaporates. What happens to the molecules or atoms that compose these substances during such changes? The answer depends on the type of change. Changes that alter only state or appearance, but not composition, are **physical changes**. The atoms or molecules that compose a substance *do not change* their identity during a physical change. For example, when water boils, it changes its state from a liquid to a gas, but the gas remains composed of water molecules, so this is a physical change (Figure 1.6▲).

A physical change results in a different form of the same substance, while a chemical change results in a completely different substance.

In contrast, changes that alter the composition of matter are **chemical changes**. During a chemical change, atoms rearrange, transforming the original substances into different substances. For example, the rusting of iron is a chemical change. The atoms that compose iron (iron atoms) combine with oxygen molecules from air to form iron oxide, the orange substance we call rust (Figure 1.7◀). Figure 1.8▶ illustrates other examples of physical and chemical changes.

Physical and chemical changes are manifestations of physical and chemical properties. A **physical property** is a property that a substance displays without changing its composition, whereas a **chemical property** is a property that a substance displays only by changing its composition via a chemical change. The smell of gasoline is a physical property—gasoline does not change its composition when it exhibits its odor. The flammability of gasoline, in contrast, is a chemical property—gasoline does change its composition when it burns, turning into completely new substances (primarily carbon dioxide and water). Physical properties include odor, taste, color, appearance, melting point, boiling point, and density. Chemical properties include corrosiveness, flammability, acidity, toxicity, and other such characteristics.

The differences between physical and chemical changes are not always apparent. Only chemical examination can confirm whether a particular change is physical or chemical. In many cases, however, we can identify chemical and physical changes based on what we know about the changes. Changes in the state of matter, such as melting or boiling, or changes in the physical condition of matter, such as those that result from cutting or crushing, are typically physical changes. Changes involving chemical reactions—often evidenced by temperature or color changes—are chemical changes.



▲ **FIGURE 1.7 Rusting, a Chemical Change** When iron rusts, the iron atoms combine with oxygen atoms to form a different chemical substance, the compound iron oxide. Rusting is a chemical change, and the tendency of iron to rust is a chemical property. A more detailed exploration of this reaction can be found in Section 20.9.

EXAMPLE 1.1 Physical and Chemical Changes and Properties

Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

(a) the evaporation of rubbing alcohol
(b) the burning of lamp oil
(c) the bleaching of hair with hydrogen peroxide
(d) the formation of frost on a cold night

SOLUTION

(a) When rubbing alcohol evaporates, it changes from liquid to gas, but it remains alcohol—this is a physical change. The volatility (the ability to evaporate easily) of alcohol is therefore a physical property.

(b) Lamp oil burns because it reacts with oxygen in air to form carbon dioxide and water—this is a chemical change. The flammability of lamp oil is therefore a chemical property.

(c) Applying hydrogen peroxide to hair changes pigment molecules in hair that give it color—this is a chemical change. The susceptibility of hair to bleaching is therefore a chemical property.

(d) Frost forms on a cold night because water vapor in air changes its state to form solid ice—this is a physical change. The temperature at which water freezes is therefore a physical property.

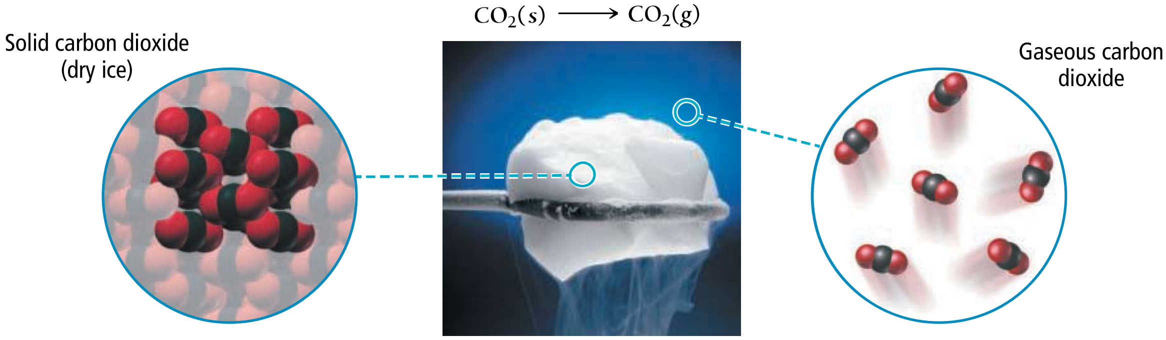
Answers to For Practice and For More Practice problems can be found in Appendix IV.

FOR PRACTICE 1.1 Determine whether each change is physical or chemical. What kind of property (chemical or physical) is demonstrated in each case?

(a) A copper wire is hammered flat.
(b) A nickel dissolves in acid to form a blue-green solution.
(c) Dry ice sublimes without melting.
(d) A match ignites when struck on a flint.

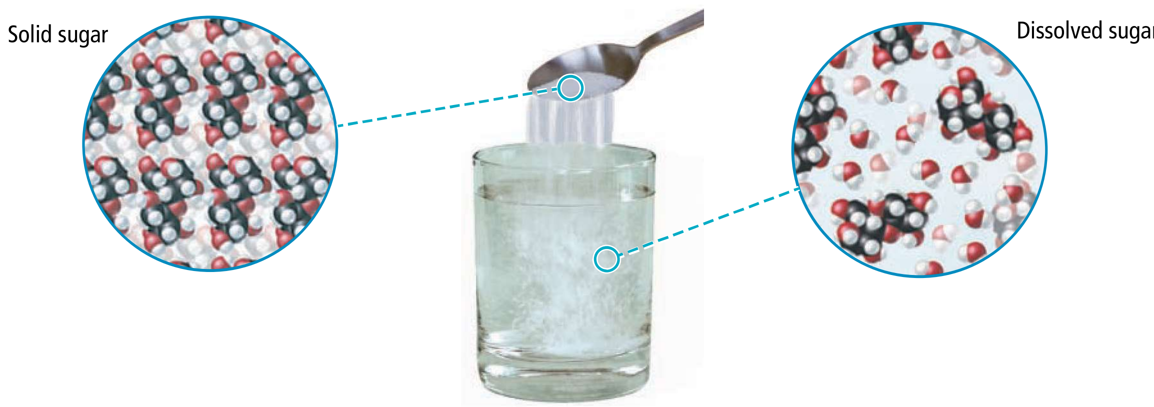
Physical Change versus Chemical Change

$$CO_2(s) \longrightarrow CO_2(g)$$



(a) Dry ice subliming. Chemical composition unaltered: **Physical change**

$$C_{12}H_{22}O_{11}(s) \longrightarrow C_{12}H_{22}O_{11}(aq)$$



(b) Sugar dissolving. Chemical composition unaltered: **Physical change**

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

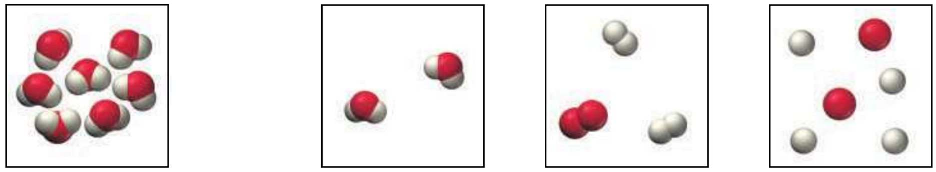
(c) Propane gas burning. Chemical composition altered: **Chemical change**

▲ **FIGURE 1.8 Physical and Chemical Changes** **(a)** The sublimation (the state change from a solid to a gas) of dry ice (solid CO_2) is a physical change. **(b)** The dissolution of sugar is a physical change. **(c)** The burning of propane is a chemical change.

ANSWER **NOW!**

1.3 Cc Conceptual Connection

CHEMICAL AND PHYSICAL CHANGES The diagram on the left represents liquid water molecules in a pan. Which of the three diagrams (a, b, or c) best represents the water molecules after they have been vaporized by boiling?

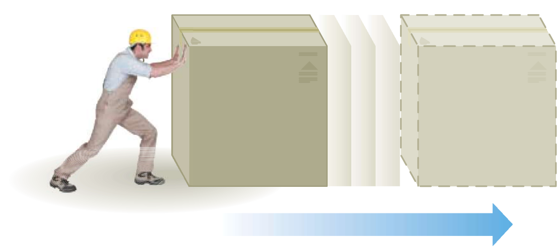


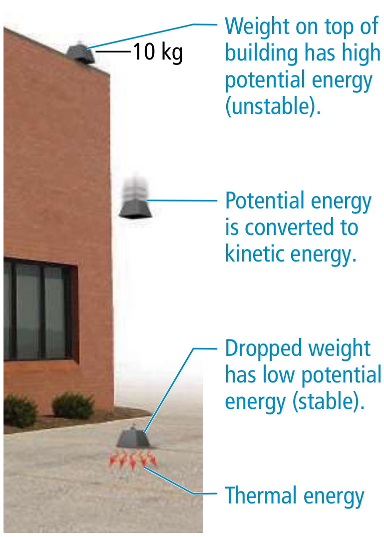
(a) **(b)** **(c)**

1.5 Energy: A Fundamental Part of Physical and Chemical Change

The physical and chemical changes discussed in Section 1.4 are usually accompanied by energy changes. For example, when water evaporates from your skin (a physical change), the water molecules absorb energy from your body, making you feel cooler. When you burn natural gas on the stove (a chemical change), energy is released, heating the food you are cooking. Understanding the physical and chemical changes of matter—that is, understanding chemistry—requires that you understand energy changes and energy flow.

The scientific definition of **energy** is *the capacity to do work*. **Work** is defined as the action of a force through a distance. For instance, when you push a box across the floor or pedal your bicycle across the street, you have done work.





▲ **FIGURE 1.9 Energy Conversions** Gravitational potential energy is converted into kinetic energy when the weight is dropped. The kinetic energy is converted mostly to thermal energy when the weight strikes the ground.

The *total energy* of an object is a sum of its **kinetic energy** (the energy associated with its motion) and its **potential energy** (the energy associated with its position or composition). For example, a weight held several meters above the ground has potential energy due to its position within Earth's gravitational field (Figure 1.9◀). If you drop the weight, it accelerates, and its potential energy is converted to kinetic energy. When the weight hits the ground, its kinetic energy is converted primarily to **thermal energy**, the energy associated with the temperature of an object. Thermal energy is actually a type of kinetic energy because it is associated with the motion of the individual atoms or molecules that make up an object. When the weight hits the ground, its kinetic energy is essentially transferred to the atoms and molecules that compose the ground, raising the temperature of the ground ever so slightly.

The first principle to note about how energy changes as the weight falls to the ground is that *energy is neither created nor destroyed*. The potential energy of the weight becomes kinetic energy as the weight accelerates toward the ground. The kinetic energy then becomes thermal energy when the weight hits the ground. The total amount of thermal energy that is released through the process is exactly equal to the initial potential energy of the weight. The idea that energy is neither created nor destroyed is known as the **law of conservation of energy**. Although energy can change from one type into another, and although it can flow from one object to another, the *total quantity* of energy does not change—it remains constant.

In Chapter 21 we will discuss how energy conservation is actually part of a more general law that allows for the interconvertibility of mass and energy.

The second principle to note about the raised weight and its fall is *the tendency of systems with high potential energy to change in a way that lowers their potential energy*. For this reason, objects or systems with high potential energy tend to be *unstable*. The weight lifted several meters from the ground is unstable because it contains a significant amount of potential energy. Unless restrained, the weight will naturally fall, lowering its potential energy (due to its position in Earth's gravitational field). We can harness some of the raised weight's potential energy to do work. For example, we can attach the weight to a rope that turns a paddle wheel or spins a drill as the weight falls. After the weight falls to the ground, it contains less potential energy—it has become more *stable*.

▲ **FIGURE 1.10 Using Chemical Energy to Do Work** The compounds produced when gasoline burns have less chemical potential energy than the gasoline molecules.

Some chemical substances are like a raised weight. For example, the molecules that compose gasoline have a relatively high potential energy—energy is concentrated in them just as energy is concentrated in the raised weight. The molecules in the gasoline tend to undergo chemical changes (specifically combustion) that lower the molecules' potential energy. As the energy of the molecules is released, some of it can be harnessed to do work, such as moving a car forward (Figure 1.10▲). The molecules that result from the chemical change have less potential energy than the original molecules in gasoline and are more stable.

Chemical potential energy, such as the energy contained in the molecules that compose gasoline, arises primarily from electrostatic forces between the electrically charged particles (protons and electrons) that compose atoms and molecules. We will learn more about those particles, as well as the properties of electrical charge, in Chapter 2, but for now, know that molecules contain specific, usually complex, arrangements of these charged particles. Some of these arrangements—such as the one within the molecules that compose gasoline—have a much higher potential energy than others. When gasoline undergoes combustion, the arrangement of these particles changes, creating molecules with much lower potential energy and transferring a great deal of energy (mostly in the form of heat) to the surroundings.

Summarizing Energy:

- Energy is always conserved in a physical or chemical change; it is neither created nor destroyed.
- Systems with high potential energy tend to change in ways that lower their potential energy, transferring energy to the surroundings.

ENERGY What type of energy is chemical energy?

(a) kinetic energy
(b) thermal energy
(c) potential energy

1.4 Cc Conceptual Connection

ANSWER **NOW!**

1.6 The Units of Measurement

WATCH **NOW!**

KEY CONCEPT VIDEO 1.6 *Units and Significant Figures*

In 1999, NASA lost the $125 million *Mars Climate Orbiter*. The chairman of the commission that investigated the disaster concluded, "The root cause of the loss of the spacecraft was a failed translation of English units into metric units." As a result, the orbiter—which was supposed to monitor weather on Mars—descended too far into the Martian atmosphere and burned up. In chemistry as in space exploration, **units**—standard quantities used to specify measurements—are critical. If we get them wrong, the consequences can be disastrous.

▲ The $125 million *Mars Climate Orbiter* was lost in 1999 because two groups of engineers used different units.

TABLE 1.1 ■ SI Base Units

Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

The two most common unit systems are the **metric system**, used in most of the world, and the **English system**, used in the United States. Scientists use the **International System of Units (SI)**, which is based on the metric system.

The abbreviation *SI* comes from the French, *Système International d'Unités*.

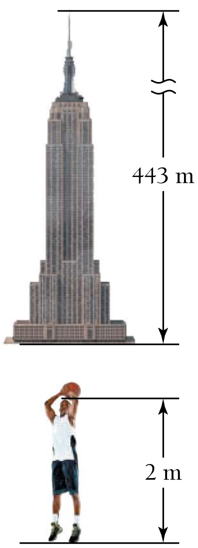
Standard Units

Table 1.1 shows the standard SI base units. In this chapter, we focus on the first four of these units: the *meter*, the standard unit of length; the *kilogram*, the standard unit of mass; the *second*, the standard unit of time; and the *kelvin*, the standard unit of temperature.

The Meter: A Measure of Length

The **meter (m)** is slightly longer than a yard (1 yard is 36 inches, while 1 meter is 39.37 inches).

Thus, a 100-yard (yd) football field measures only 91.4 m. The meter was originally defined as 1/10,000,000 of the distance from the equator to the North Pole (through Paris). The International Bureau of Weights and Measures now defines it more precisely as the distance light travels through a vacuum in a certain period of time, 1/299,792,458 second. A tall human is about 2 m tall, and the Empire State Building stands 443 m tall (including its mast).



▲ The Empire State Building is 443 m tall. A basketball player stands about 2 m tall.

The Kilogram: A Measure of Mass

The **kilogram (kg)**, defined as the mass of a metal cylinder kept at the International Bureau of Weights and Measures at Sèvres, France, is a measure of *mass*, a quantity different from *weight*. The **mass** of an object is a measure of the quantity of matter within it, while the weight of an object is a measure of the *gravitational pull* on its matter. If you could weigh yourself on the moon, for example, its weaker gravity would pull on you with less force than does Earth's gravity, resulting in a lower weight. A 130-pound (lb) person on Earth would weigh only 21.5 lb on the moon. However, a person's mass—the quantity of matter in his or her body—remains the same on every planet. One kilogram of mass is the equivalent of 2.205 lb of weight on Earth, so expressed in kilograms, a 130-lb person has a mass of approximately 59 kg and this book has a mass of about 2.5 kg. A second common unit of mass is the gram (g). One gram is 1/1000 kg. A nickel (5¢) has a mass of about 5 g.

▲ A nickel (5 cents) weighs about 5 g.

The Second: A Measure of Time

If you live in the United States, the **second (s)** is perhaps the SI unit most familiar to you. The International Bureau of Weights and Measures originally defined the second in terms of the day and the year, but a second is now defined more precisely as the duration of 9,192,631,770 periods of the radiation emitted from a certain transition in a cesium-133 atom. (We discuss transitions and the emission of radiation by atoms in Chapter 8.)

Scientists measure time on a large range of scales. The human heart beats about once every second, the age of the universe is estimated to be about 4.32×10^{17} s (13.7 billion years), and some molecular bonds break or form in time periods as short as 1×10^{-15} s.

The Kelvin: A Measure of Temperature

The **kelvin (K)** is the SI unit of **temperature**. The temperature of a sample of matter is a measure of the average kinetic energy—the energy due to motion—of the atoms or molecules that compose the matter. The molecules in a *hot* glass of water are, on average, moving faster than the molecules in a *cold* glass of water. Temperature is a measure of this molecular motion.

Temperature also determines the direction of thermal energy transfer, what we commonly call *heat*. Thermal energy transfers from hot objects to cold ones. For example, when you touch another person's warm hand (and yours is cold), thermal energy flows *from his or her hand to yours*, making your hand feel warmer. However, if you touch an ice cube, thermal energy flows *out of your hand* to the ice, cooling your hand (and possibly melting some of the ice cube).

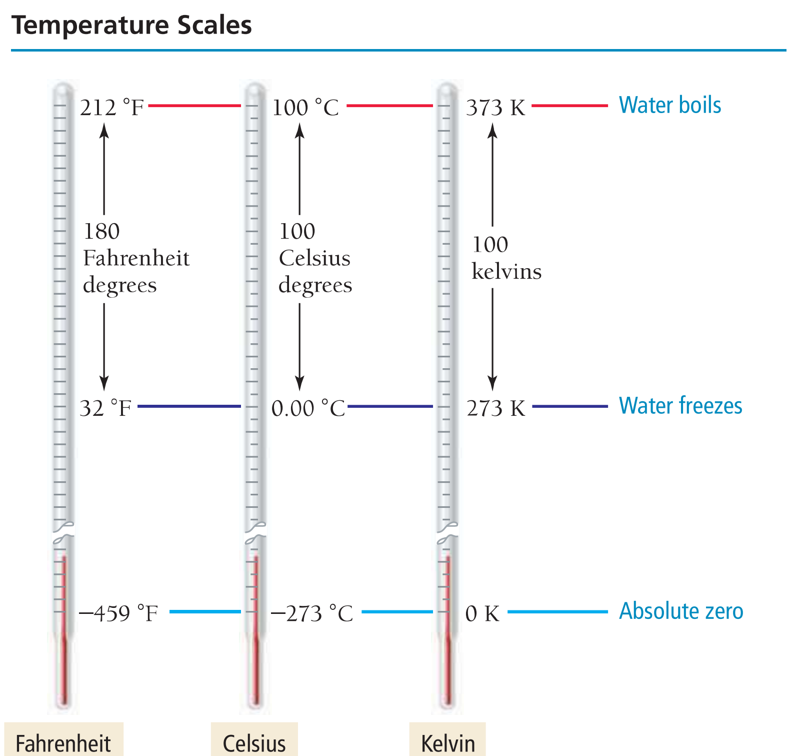
Figure 1.11▼ shows the three common temperature scales. The most common in the United States is the **Fahrenheit (°F) scale**, shown on the left in Figure 1.11. On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F at sea level. Room temperature is approximately 72 °F. The Fahrenheit scale was originally determined by assigning 0 °F to the freezing point of a concentrated saltwater solution and 96 °F to normal body temperature.

Normal body temperature on the modern Farhenheit scale is 98.6 °F.

Scientists and citizens of most countries other than the United States typically use the **Celsius (°C) scale**, shown in the middle in Figure 1.11. On this scale, pure water freezes at 0 °C and boils at 100 °C (at sea level). Room temperature is approximately 22 °C. The Fahrenheit scale and the Celsius scale differ both in the size of their respective degrees and the temperature each designates as "zero." Both the Fahrenheit and Celsius scales allow for negative temperatures.

The SI unit for temperature is the kelvin, shown in Figure 1.11. The **Kelvin scale** (sometimes also called the *absolute scale*) avoids negative temperatures by assigning 0 K to the coldest temperature possible, absolute zero. Absolute zero (–273.15 °C or –459 °F) is the temperature at which molecular motion virtually stops. Lower temperatures do not exist. The size of the kelvin is identical to that of the Celsius degree; the only difference is

Molecular motion does not *completely* stop at absolute zero because of the uncertainty principle in quantum mechanics, which we discuss in Chapter 8.



◀ **FIGURE 1.11 Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales** The Fahrenheit degree is five-ninths the size of the Celsius degree and the kelvin. The zero point of the Kelvin scale is absolute zero (the lowest possible temperature), whereas the zero point of the Celsius scale is the freezing point of water.

The Celsius Temperature Scale

0 °C – Water freezes

10 °C – Brisk fall day

22 °C – Room temperature

45 °C – Summer day in Death Valley

the temperature that each scale designates as zero. We can convert between the temperature scales with these formulas:

$$°\text{C} = \frac{(°\text{F} - 32)}{1.8}$$

$$\text{K} = °\text{C} + 273.15$$

Note that we give Kelvin temperatures in kelvins (*not* "degrees Kelvin") or K (*not* °K).

ANSWER **NOW!**

1.5 Cc Conceptual Connection

TEMPERATURE SCALES Which temperature scale has no negative temperatures?

(a) Kelvin
(b) Celsius
(c) Fahrenheit

Throughout this book, we provide examples worked out in formats that are designed to help you develop problem-solving skills. The most common format uses two columns to guide you through the worked example. The left column describes the thought processes and steps used in solving the problem, while the right column shows the implementation. Example 1.2 follows this two-column format.

EXAMPLE 1.2 Converting between Temperature Scales

A sick child has a temperature of 40.00 °C. What is the child's temperature in **(a)** K and **(b)** °F?

SOLUTION

(a) Begin by finding the equation that relates the quantity that is given (°C) and the quantity you are trying to find (K).	$\text{K} = °\text{C} + 273.15$
Since this equation gives the temperature in K directly, substitute in the correct value for the temperature in °C and calculate the answer.	$\text{K} = °\text{C} + 273.15$ $\text{K} = 40.00 + 273.15 = 313.15\ \text{K}$
(b) To convert from °C to °F, first find the equation that relates these two quantities.	$°\text{C} = \frac{(°\text{F} - 32)}{1.8}$
Since this equation expresses °C in terms of °F, solve the equation for °F.	$°\text{C} = \frac{(°\text{F} - 32)}{1.8}$ $1.8(°\text{C}) = (°\text{F} - 32)$ $°\text{F} = 1.8(°\text{C}) + 32$
Now substitute °C into the equation and calculate the answer. *Note: The number of digits reported in this answer follows significant figure conventions, covered in Section 1.7.*	$°\text{F} = 1.8(°\text{C}) + 32$ $°\text{F} = 1.8(40.00\ °\text{C}) + 32 = 104.00\ °\text{F}$

FOR PRACTICE 1.2 Gallium is a solid metal at room temperature, but it will melt to a liquid in your hand. The melting point of gallium is 85.6 °F. What is this temperature on **(a)** the Celsius scale and **(b)** the Kelvin scale?

Prefix Multipliers

Scientific notation (see Appendix IA) allows us to express very large or very small quantities in a compact manner by using exponents. For example, the diameter of a hydrogen atom is 1.06×10^{-10} m. The International System of Units uses the **prefix multipliers** listed in Table 1.2 with the standard units. These multipliers change the value of the unit by powers of 10 (just like an exponent does in scientific notation). For example, the kilometer has the prefix *kilo* meaning 1000 or 10^3. Therefore,

$$1 \text{ kilometer} = 1000 \text{ meters} = 10^3 \text{ meters}$$

Similarly, the millimeter has the prefix *milli* meaning 0.001 or 10^{-3}.

$$1 \text{ millimeter} = 0.001 \text{ meters} = 10^{-3} \text{ meters}$$

TABLE 1.2 ■ SI Prefix Multipliers

Prefix	Symbol	Multiplier	
exa	E	1,000,000,000,000,000,000	(10^{18})
peta	P	1,000,000,000,000,000	(10^{15})
tera	T	1,000,000,000,000	(10^{12})
giga	G	1,000,000,000	(10^{9})
mega	M	1,000,000	(10^{6})
kilo	k	1000	(10^{3})
deci	d	0.1	(10^{-1})
centi	c	0.01	(10^{-2})
milli	m	0.001	(10^{-3})
micro	μ	0.000001	(10^{-6})
nano	n	0.000000001	(10^{-9})
pico	p	0.000000000001	(10^{-12})
femto	f	0.000000000000001	(10^{-15})
atto	a	0.000000000000000001	(10^{-18})

When reporting a measurement, choose a prefix multiplier close to the size of the quantity you are measuring. For example, to state the diameter of a hydrogen atom, which is 1.06×10^{-10} m, use picometers (106 pm) or nanometers (0.106 nm) rather than micrometers or millimeters. Choose the prefix multiplier that is most convenient for a particular number.

PREFIX MULTIPLIERS Which prefix multiplier is most appropriate for reporting a measurement of 5.57×10^{-5} m?

(a) mega
(b) milli
(c) micro
(d) kilo

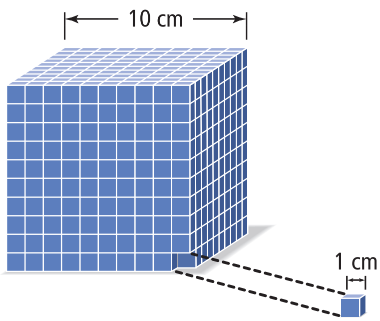
1.6 Cc Conceptual Connection

ANSWER NOW!

Derived Units: Volume and Density

A **derived unit** is a combination of other units. For example, the SI unit for speed is meters per second (m/s), a derived unit. Notice that this unit is formed from two other SI units—meters and seconds—put together. You are probably more familiar with speed measured in miles/hour or kilometers/hour—these are also examples of derived units. Two other common derived units are those for volume (SI base unit is m^3) and density (SI base unit is kg/m^3).

Relationship between Length and Volume



▲ **FIGURE 1.12 The Relationship between Length and Volume** A cube with a 10-cm edge has a volume of $(10\ \text{cm})^3$ or $1000\ \text{cm}^3$, and a cube with a 100-cm edge has a volume of $(100\ \text{cm})^3 = 1{,}000{,}000\ \text{cm}^3$.

Volume

Volume is a measure of space. Any unit of length, when cubed (raised to the third power), becomes a unit of volume. The cubic meter (m^3), cubic centimeter (cm^3), and cubic millimeter (mm^3) are all units of volume. The cubic nature of volume is not always intuitive, and studies have shown that our brains are not naturally wired to process abstract concepts such as volume. For example, consider this question: how many small cubes measuring 1 cm on each side are required to construct a large cube measuring 10 cm on a side?

The answer to this question, as you can see by carefully examining the unit cube in Figure 1.12◀, is 1000 small cubes. When we go from a linear, one-dimensional distance to three-dimensional volume, we must raise both the linear dimension *and* its unit to the third power (not multiply by 3). Thus, the volume of a cube is equal to the length of its edge cubed.

$$\text{volume of cube} = (\text{edge length})^3$$

Other common units of volume in chemistry are the **liter (L)** and the **milliliter (mL)**. One milliliter (10^{-3} L) is equal to 1 cm^3. A gallon of gasoline contains 3.785 L. Table 1.3 lists some common units—for volume and other quantities—and their equivalents.

TABLE 1.3 ■ Some Common Units and Their Equivalents

Length	Mass	Volume
1 kilometer (km) = 0.6214 mile (mi)	1 kilogram (kg) = 2.205 pounds (lb)	1 liter (L) = 1000 mL = 1000 cm^3
1 meter (m) = 39.37 inches (in) = 1.094 yards (yd)	1 pound (lb) = 453.59 grams (g)	1 liter (L) = 1.057 quarts (qt)
1 foot (ft) = 30.48 centimeters (cm)	1 ounce (oz) = 28.35 grams (g)	1 U.S. gallon (gal) = 3.785 liters (L)
1 inch (in) = 2.54 centimeters (cm) (exact)		

Density

An old riddle asks, "Which weighs more, a ton of bricks or a ton of feathers?" The answer is neither—they both weigh the same (1 ton). If you answered bricks, you confused weight with density. The **density *(d)*** of a substance is the ratio of its mass (m) to its volume (V).

The m in the equation for density is in italic type, meaning that it stands for mass rather than for meters. In general, the symbols for units such as meters (m), seconds (s), or kelvins (K) appear in regular type while those for variables such as mass (m), volume (V), and time (t) appear in italics.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad d = \frac{m}{V}$$

Density is a characteristic physical property of substances (see Table 1.4) that depends on temperature. Density is an example of an **intensive property**, one that is *independent* of the amount of the substance. The density of aluminum, for example, is the same whether you have a gram or a kilogram. Intensive properties are often used to identify substances because these properties depend only on the type of substance, not on the amount of it. For example, from Table 1.4 you can see that pure gold has a density of 19.3 g/cm^3. One way to determine whether a substance is pure gold is to determine its density and compare it to 19.3 g/cm^3. Mass, in contrast, is an **extensive property**, one that depends on the amount of the substance. If you know only the mass of a sample of gold, that information alone will not allow you to identify it as gold.

The units of density are those of mass divided by volume. Although the SI-derived unit for density is kg/m^3, the density of liquids and solids is most often expressed in g/cm^3 or g/mL. (Remember that cm^3 and mL are equivalent units: 1 cm^3 = 1 mL.) Aluminum is among the least dense metals with a density of 2.7 g/cm^3, while platinum is one of the densest metals with a density of 21.4 g/cm^3.

Calculating Density

We can calculate the density of a substance by dividing the mass of a given amount of the substance by its volume. For example, suppose a small nugget we suspect to be gold has a mass of 22.5 g and a volume of 2.38 cm^3. To find its density, we divide the mass by the volume:

$$d = \frac{m}{V} = \frac{22.5\ \text{g}}{2.38\ \text{cm}^3} = 9.45\ \text{g/cm}^3$$

In this case, the density reveals that the nugget is not pure gold because the density of gold is 19.3 g/cm^3.

TABLE 1.4 ■ The Density of Some Common Substances at 20 °C

Substance	Density (g/cm^3)
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

EXAMPLE 1.3 Calculating Density

A man receives a platinum ring from his fiancée. Before the wedding, he notices that the ring feels a little light for its size, and so he decides to determine its density. He places the ring on a balance and finds that it has a mass of 3.15 g. He then finds that the ring displaces 0.233 cm^3 of water. Is the ring made of platinum? (Note: The volume of irregularly shaped objects is often measured by the displacement of water. To use this method, the object is placed in water and the change in volume of the water is measured. This increase in the total volume represents the volume of water *displaced* by the object and is equal to the volume of the object.)

Set up the problem by writing the important information that is *given* as well as the information that you are asked to *find*. In this case, we need to find the density of the ring and compare it to that of platinum. *Note: In Section 1.8, we discuss this standardized way of setting up problems.*	**GIVEN:** $m = 3.15\ \text{g}$ $V = 0.233\ \text{cm}^3$ **FIND:** Density in g/cm^3
Next, write down the equation that defines density.	**EQUATION** $d = \frac{m}{V}$
Solve the problem by substituting the correct values of mass and volume into the expression for density.	**SOLUTION** $d = \frac{m}{V} = \frac{3.15\ \text{g}}{0.233\ \text{cm}^3} = 13.5\ \text{g/cm}^3$

The density of the ring is much too low to be platinum (platinum density is 21.4 g/cm^3), and the ring is therefore a fake.

FOR PRACTICE 1.3 The woman in Example 1.3 is shocked that the ring is fake and returns it. She buys a new ring that has a mass of 4.53 g and a volume of 0.212 cm^3. Is the new ring genuine?

FOR MORE PRACTICE 1.3 A metal cube has an edge that is 11.4 mm long and a mass of 6.67 g. Calculate the density of the metal and use Table 1.4 to determine the likely identity of the metal.

DENSITY The density of copper decreases as temperature increases (as does the density of most substances). Which change occurs in a sample of copper when it is warmed from room temperature to 95 °C?

(a) The sample becomes lighter.
(b) The sample becomes heavier.
(c) The sample expands.
(d) The sample contracts.

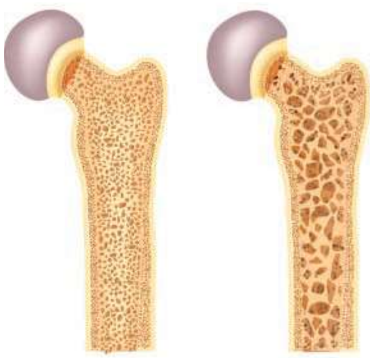
1.7 Cc Conceptual Connection

ANSWER NOW!

CHEMISTRY AND MEDICINE | Bone Density

Osteoporosis—which means *porous bone*—is a condition in which bone density becomes low. The healthy bones of a young adult have a density of about 1.0 g/cm^3. Patients suffering from osteoporosis, however, can have bone densities as low as 0.22 g/cm^3. These low densities indicate that the bones have deteriorated and weakened, resulting in increased susceptibility to fractures, especially hip fractures. Patients suffering from osteoporosis can also experience height loss and disfiguration such as dowager's hump, a condition in which the patient becomes hunched over due to compression of the vertebrae. Osteoporosis is most common in postmenopausal women, but it can also occur in people (including men) who have certain diseases, such as insulin-dependent diabetes, or who take certain medications, such as prednisone. Osteoporosis is usually diagnosed and monitored with hip X-rays. Low-density bones absorb less of the X-rays than do high-density bones, producing characteristic differences in the X-ray image. Treatments for osteoporosis include calcium and vitamin D supplements, drugs that prevent bone weakening, exercise and strength training, and, in extreme cases, hip-replacement surgery.

QUESTION *Suppose you find a large animal bone in the woods, too large to fit in a beaker or flask. How might you approximate its density?*



◀ Top: Severe osteoporosis can necessitate surgery to implant an artificial hip joint, seen in this X-ray image. Bottom: Views of the bone matrix in a normal bone (left) and one weakened by osteoporosis (right).

1.7 The Reliability of a Measurement

Carbon monoxide is a colorless gas emitted by motor vehicles and found in polluted air. The table shown here lists carbon monoxide concentrations in Los Angeles County as reported by the U.S. Environmental Protection Agency (EPA) over the period 1990–2017:

Year	Carbon Monoxide Concentration (ppm)*
1990	9.9
1995	9.0
2000	7.4
2005	2.8
2010	1.9
2017	1.1

*Second maximum, 8 hour average; ppm = parts per million (Pasadena Site 06-037-2005)

The first thing you should notice about these values is that they decrease over time. For this decrease, we can thank the Clean Air Act and its amendments, which have resulted in more efficient engines and specially blended fuels and consequently in cleaner air in all major U.S. cities over the last 30 years. The second thing you might notice is the number of digits to which the measurements are reported. The number of digits in a reported measurement indicates the certainty associated with that measurement. A less certain measurement of carbon monoxide levels might be reported as follows:

Year	Carbon Monoxide Concentration (ppm)
1990	10
1995	9
2000	7
2005	3
2010	2
2017	1

Notice that the first set of data is reported to the nearest 0.1 ppm, while the second set is reported to the nearest 1 ppm. Scientists report measured quantities in an agreed-upon standard way. The number of reported digits reflects the certainty in the measurement: more digits, more certainty; fewer digits, less certainty. Numbers are usually written so that the uncertainty is in the last reported digit. (We assume that uncertainty to be ± 1 in the last digit unless otherwise indicated.) By reporting the 2010 carbon monoxide concentration as 1.9 ppm, the scientists mean 1.9 ± 0.1 ppm. The carbon monoxide concentration is between 1.8 and 2.0 ppm—it might be 2.0 ppm, for example, but it could not be 3.0 ppm. In contrast, if the reported value was 2 ppm (as in the second set of measurements), this would mean 2 ± 1 ppm, or between 1 and 3 ppm. In general,

Scientific measurements are reported so that every digit is certain except the last, which is estimated.

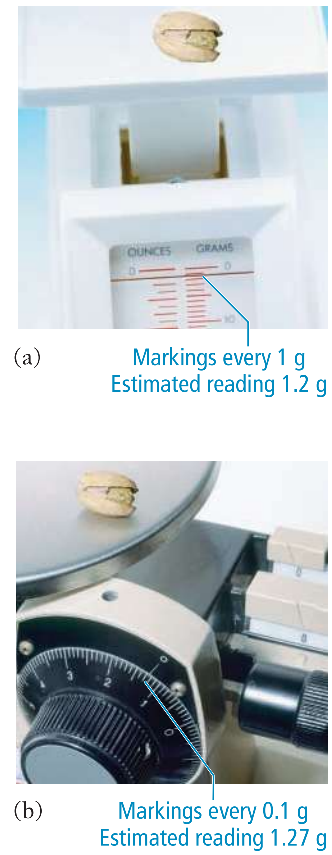
For example, consider the following reported number:

$$\underbrace{5.21}_{\text{certain}}\underbrace{3}_{\text{estimated}}$$

The first three digits are certain; the last digit is estimated.

The number of digits reported in a measurement depends on the measuring device. Consider weighing a pistachio nut on two different balances (Figure 1.13▶). The balance on the top has marks every 1 g, while the balance on the bottom has marks every 0.1 g. For the balance on the top, we mentally divide the space between the 1- and 2-g marks into ten equal spaces and estimate that the pointer is at about 1.2 g. We then write the measurement as 1.2 g, indicating that we are sure of the "1" but we have estimated the ".2." The balance on the bottom, with marks every *tenth* of a gram, requires that we write the result with more digits. The pointer is between the 1.2-g mark and the 1.3-g mark. We again divide the space between the two marks into ten equal spaces and estimate the third digit. For the figure shown, we report 1.27 g.

Estimation in Weighing



▲ **FIGURE 1.13 Estimation in Weighing** **(a)** This scale has markings every 1 g, so we mentally divide the space into ten equal spaces to estimate the last digit. This reading is 1.2 g. **(b)** Because this balance has markings every 0.1 g, we estimate to the hundredths place. This reading is 1.27 g.

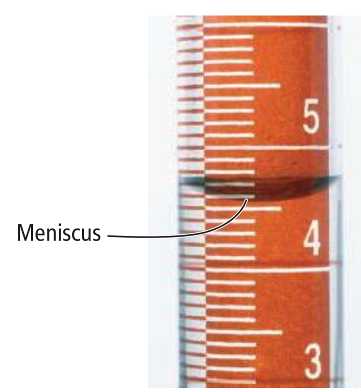
EXAMPLE 1.4 Reporting the Correct Number of Digits

The graduated cylinder shown in the right margin has markings every 0.1 mL. Report the volume (which is read at the bottom of the meniscus) to the correct number of digits. (Note: The meniscus is the crescent-shaped surface at the top of a column of liquid.)

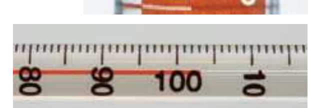
SOLUTION

Since the bottom of the meniscus is between the 4.5 and 4.6 mL markings, mentally divide the space between the markings into ten equal spaces and estimate the next digit. In this case, you should report the result as 4.57 mL.

What if you estimated a little differently and wrote 4.56 mL? In general, a one-unit difference in the last digit is acceptable because the last digit is estimated and different people might estimate it slightly differently. However, if you wrote 4.63 mL, you would have misreported the measurement.



FOR PRACTICE 1.4 Record the temperature on the thermometer shown in the right margin to the correct number of digits.



Counting Significant Figures

The precision of a measurement—which depends on the instrument used to make the measurement—must be preserved, not only when recording the measurement, but also when performing calculations that use the measurement. We can accomplish the preservation of this precision by using *significant figures*. In any reported measurement, the non–place-holding digits—those that are not simply marking the decimal place—are called **significant figures** (or **significant digits**). *The greater the number of significant figures, the greater the certainty of the measurement*. For example, the number 23.5 has three significant figures while the number 23.56 has four. To determine the number of significant figures in a number containing zeroes, we distinguish between zeroes that are significant and those that simply mark the decimal place. For example, in the number 0.0008, the leading zeroes (zeroes to the left of the first nonzero digit) mark the decimal place but *do not* add to the certainty of the measurement and are therefore not significant; this number has only one significant figure. In contrast, the trailing zeroes (zeroes at the end of a number) in the number 0.000800 *do add* to the certainty of the measurement and are therefore counted as significant; this number has three significant figures.

HOW TO: *Determine the Number of Significant Figures in a Given Value*

Significant Figure Rules	**Examples**	
1. All nonzero digits are significant.	28.03	0.0540
2. Interior zeroes (zeroes between two nonzero digits) are significant.	408	7.0301
3. Leading zeroes (zeroes to the left of the first nonzero digit) are not significant. They only serve to locate the decimal point.	0.0032	0.00006
	not significant	
4. Trailing zeroes (zeroes at the end of a number) are categorized as follows:		
• Trailing zeroes after a decimal point are always significant.	45.000	3.5600
• Trailing zeroes before a decimal point (and after a nonzero number) are always significant.	140.00	2500.55
• Trailing zeroes before an *implied* decimal point are ambiguous and should be avoided by using scientific notation.	1200	ambiguous
	1.2×10^3	2 significant figures
	1.20×10^3	3 significant figures
	1.200×10^3	4 significant figures
• Some textbooks put a decimal point after one or more trailing zeroes if the zeroes are to be considered significant. We avoid that practice in this book, but you should be aware of it.	1200.	4 significant figures (common in some textbooks)

Exact Numbers

Exact numbers have no uncertainty and thus do not limit the number of significant figures in any calculation. We can regard an exact number as having an unlimited number of significant figures. Exact numbers originate from three sources:

- From the accurate counting of discrete objects. For example, 3 atoms means 3.00000. . . atoms.
- From defined quantities, such as the number of centimeters in 1 m. Because 100 cm is defined as 1 m,

$$100 \text{ cm} = 1 \text{ m means } 100.00000\ldots \text{ cm} = 1.0000000\ldots \text{ m}$$

- From integral numbers that are part of an equation. For example, in the equation $radius = \frac{diameter}{2}$, the number 2 is exact and therefore has an unlimited number of significant figures.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 1.5

EXAMPLE 1.5 Determining the Number of Significant Figures in a Number

How many significant figures are in each number?

(a) 0.04450 m **(b)** 5.0003 km **(c)** 10 dm = 1 m **(d)** 1.000×10^5 s **(e)** 0.00002 mm **(f)** 10,000 m

SOLUTION

(a) 0.04450 m	*Four significant figures*. The two 4s and the 5 are significant (Rule 1). The trailing zero is after a decimal point and is therefore significant (Rule 4). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(b) 5.0003 km	*Five significant figures*. The 5 and 3 are significant (Rule 1), as are the three interior zeroes (Rule 2).
(c) 10 dm = 1 m	*Unlimited significant figures*. Defined quantities have an unlimited number of significant figures.
(d) 1.000×10^5 s	*Four significant figures*. The 1 is significant (Rule 1). The trailing zeroes are after a decimal point and therefore significant (Rule 4).
(e) 0.00002 mm	*One significant figure*. The 2 is significant (Rule 1). The leading zeroes only mark the decimal place and are therefore not significant (Rule 3).
(f) 10,000 m	*Ambiguous*. The 1 is significant (Rule 1), but the trailing zeroes occur before an implied decimal point and are therefore ambiguous (Rule 4). Without more information, you would assume one significant figure. It is better to write this as 1×10^5 to indicate one significant figure or as 1.0000×10^5 to indicate five (Rule 4).

FOR PRACTICE 1.5 How many significant figures are in each number?

(a) 554 km **(b)** 7 pennies **(c)** 1.01×10^5 m **(d)** 0.00099 s **(e)** 1.4500 km **(f)** 21,000 m

Significant Figures in Calculations

WATCH **NOW!**

KEY CONCEPT VIDEO 1.7

Significant Figures in Calculations

When we use measured quantities in calculations, the results of the calculation must reflect the precision of the measured quantities. We should not lose or gain precision during mathematical operations. Follow these rules when carrying significant figures through calculations.

HOW TO: *Determine Significant Figures in Calculated Quantities*

Rules for Calculations

1. In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

2. In addition or subtraction, the result carries the same number of decimal places as the quantity with the fewest decimal places.

Examples

$$\underset{\text{(4 sig. figures)}}{1.052} \times \underset{\text{(5 sig. figures)}}{12.504} \times \underset{\text{(2 sig. figures)}}{0.53} = 6.7208 = \underset{\text{(2 sig. figures)}}{6.7}$$

$$\underset{\text{(5 sig. figures)}}{2.0035} \div \underset{\text{(3 sig. figures)}}{3.20} = 0.626094 = \underset{\text{(3 sig. figures)}}{0.626}$$

$$\begin{array}{r} 2.345 \\ 0.07 \\ 2.9975 \\ \hline 5.4125 \end{array} = 5.41 \qquad \begin{array}{r} 5.9 \\ -0.221 \\ \hline 5.679 \end{array} = 5.7$$

In addition and subtraction, it is helpful to draw a line next to the number with the fewest decimal places. This line determines the number of decimal places in the answer.

—*Continued on the next page*

Rules for Calculations	Examples
3. When rounding to the correct number of significant figures, round down if the last (or leftmost) digit dropped is four or less; round up if the last (or leftmost) digit dropped is five or more.	Rounding to two significant figures: 5.37 rounds to 5.4 5.34 rounds to 5.3 5.35 rounds to 5.4 5.349 rounds to 5.3 Notice in the last example (5.349) that only the *leftmost digit being dropped* (in this example, the 4) determines how you round the number. You ignore all the digits to the right of the leftmost digit you are dropping (in this example, you ignore the 9).
4. To avoid rounding errors in multi-step calculations, round only the final answer—do not round intermediate steps. If you write down intermediate answers, keep track of significant figures by underlining the least significant digit.	$6.78 \times 5.903 \times (5.489 - 5.01)$ $= 6.78 \times 5.903 \times 0.4\underline{7}9$ $= 19.1707$ $= 19$ underline least significant digit

A few books recommend a slightly different rounding procedure for cases in which the last digit is 5. However, the procedure presented here is consistent with electronic calculators and will be used throughout this book.

Notice that for multiplication or division, the quantity with the fewest *significant figures* determines the number of *significant figures* in the answer, but for addition and subtraction, the quantity with the fewest *decimal places* determines the number of *decimal places* in the answer. In multiplication and division, we focus on significant figures, but in addition and subtraction we focus on decimal places. When a problem involves addition or subtraction, the answer may have a different number of significant figures than the initial quantities. Keep this in mind in problems that involve both addition or subtraction and multiplication or division. For example,

$$\frac{1.002 - 0.999}{3.754} = \frac{0.003}{3.754}$$
$$= 7.99 \times 10^{-4}$$
$$= 8 \times 10^{-4}$$

The answer has only one significant figure, even though the initial numbers had three or four.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 1.6

EXAMPLE 1.6 Significant Figures in Calculations

Perform each calculation to the correct number of significant figures.

(a) $1.10 \times 0.5120 \times 4.0015 \div 3.4555$

(b)
$$\begin{array}{r} 0.355 \\ +105.1 \\ \underline{-100.5820} \end{array}$$

(c) $4.562 \times 3.99870 \div (452.6755 - 452.33)$

(d) $(14.84 \times 0.55) - 8.02$

SOLUTION

(a) Round the intermediate result (in blue) to three significant figures to reflect the three significant figures in the quantity with the fewest significant figures (1.10).	$1.10 \times 0.5120 \times 4.0015 \div 3.4555$ $= 0.65219$ $= 0.652$
(b) Round the intermediate answer (in blue) to one decimal place to reflect the quantity with the fewest decimal places (105.1). Notice that 105.1 is *not* the quantity with the fewest significant figures, but it has the fewest decimal places and therefore determines the number of decimal places in the answer.	$\begin{array}{r} 0.355 \\ +105.1 \\ \underline{-100.5820} \\ 4.8730 \end{array} = 4.9$

(c) Mark the intermediate result to two decimal places to reflect the number of decimal places in the quantity within the parentheses having the fewest number of decimal places (452.33). Round the final answer to two significant figures to reflect the two significant figures in the least precisely known quantity ($0.3\underline{4}55$).	$4.562 \times 3.99870 \div (452.6755 - 452.33)$ $= 4.562 \times 3.99870 \div 0.3\underline{4}55$ $= 52.79904$ $= 53$ 2 places of the decimal
(d) Mark the intermediate result to two significant figures to reflect the number of significant figures in the quantity within the parentheses having the fewest number of significant figures (0.55). Round the final answer to one decimal place to reflect the one decimal place in the least precisely known quantity ($8.\underline{1}62$).	$(14.84 \times 0.55) - 8.02 = 8.\underline{1}62 - 8.02$ $= 0.142$ $= 0.1$

FOR PRACTICE 1.6 Perform each calculation to the correct number of significant figures.

(a) $3.10007 \times 9.441 \times 0.0301 \div 2.31$

(b)
$$\begin{array}{r} 0.881 \\ +132.1 \\ \underline{-\ 12.02} \end{array}$$

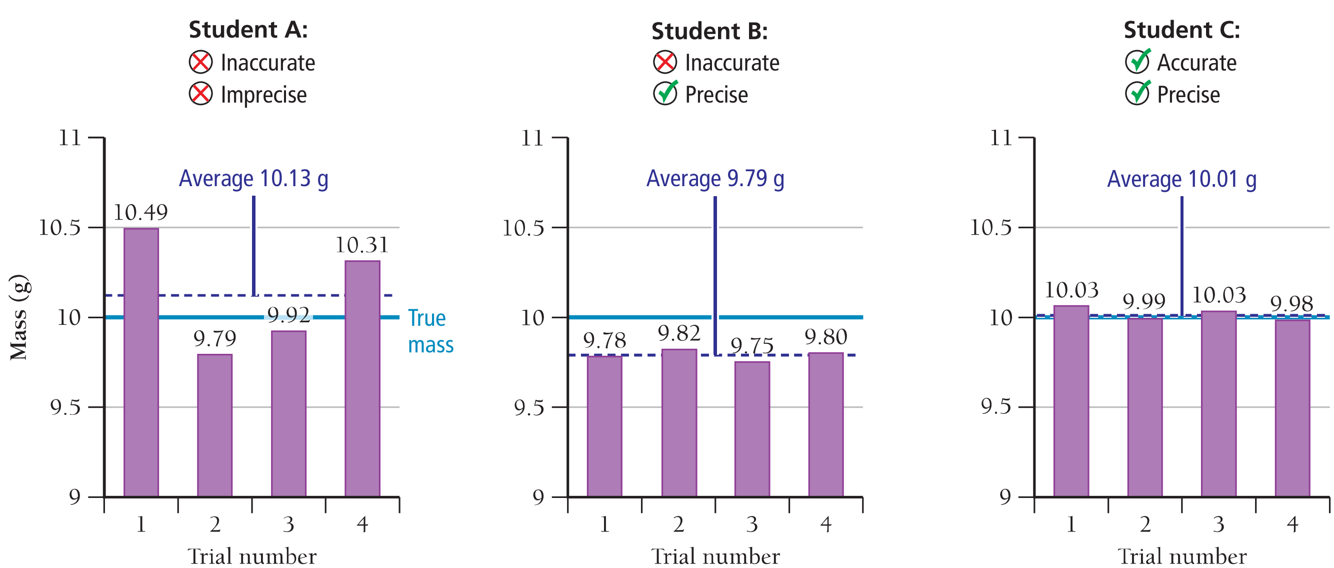
(c) $2.5110 \times 21.20 \div (44.11 + 1.223)$

(d) $(12.01 \times 0.3) + 4.811$

Precision and Accuracy

Scientists often repeat measurements several times to increase confidence in the result. We distinguish between two different kinds of certainty—accuracy and precision—associated with such measurements. **Accuracy** refers to how close the measured value is to the actual value. **Precision** refers to how close a series of measurements are to one another or how reproducible they are. A series of measurements can be precise (close to one another in value and reproducible) but not accurate (not close to the true value). Consider the results of three students who repeatedly weighed a lead block known to have a true mass of 10.00 g (indicated by the solid horizontal blue line on the graphs in the accompanying figure).

	Student A	Student B	Student C
Trial 1	10.49 g	9.78 g	10.03 g
Trial 2	9.79 g	9.82 g	9.99 g
Trial 3	9.92 g	9.75 g	10.03 g
Trial 4	10.31 g	9.80 g	9.98 g
Average	**10.13 g**	**9.79 g**	**10.01 g**



▲ Measurements are said to be precise if they are consistent with one another, but they are accurate only if they are close to the actual value.

- The results of student A are both inaccurate (not close to the true value) and imprecise (not consistent with one another). The inconsistency is the result of **random error**, error that has equal probability of being too high or too low. Almost all measurements have some degree of random error. Random error can, with enough trials, average itself out.
- The results of student B are precise (close to one another in value) but inaccurate. The inaccuracy is the result of **systematic error**, error that tends to be either too high or too low. In contrast to random error, systematic error does not average out with repeated trials. For instance, if a balance is not properly calibrated, it will systematically read too high or too low.
- The results of student C display little systematic error or random error—they are both accurate and precise.

CHEMISTRY IN YOUR DAY | Integrity in Data Gathering

Most scientists spend many hours collecting data in the laboratory. Often, the data do not turn out exactly as the scientist had expected (or hoped). A scientist may then be tempted to "fudge" the results. For example, suppose you are expecting a particular set of measurements to follow a certain pattern. After working hard over several days or weeks to make the measurements, you notice that a few of them do not quite fit the pattern that you anticipated. You might find yourself wishing that you could simply change or omit the "faulty" measurements. Altering data in this way is considered highly unethical in the scientific community and, when discovered, is punished severely.

In 2004, Dr. Hwang Woo Suk, a stem cell researcher at the Seoul National University in South Korea, published a research paper in *Science* (a highly respected research journal) claiming that he and his colleagues had cloned human embryonic stem cells. As part of his evidence, he showed photographs of the cells. The paper was hailed as an incredible breakthrough, and Dr. Hwang traveled the world lecturing on his work. *Time* magazine even included him on its "people that matter" list for 2004. Several months later, however, one of his coworkers revealed that the photographs were fraudulent. According to the coworker, the photographs came from a computer data bank of stem cell photographs, not from a cloning experiment. A university panel investigated the results and confirmed that the photographs and other data had indeed been faked. Dr. Hwang was forced to resign his prestigious post at the university.

Although not common, incidents like this do occur from time to time. They are damaging to a community that is largely built on trust. A scientist's peers (other researchers in similar fields) review all published research papers, but usually they are judging whether the data support the conclusion—they assume that the experimental measurements are authentic. The pressure to succeed sometimes leads researchers to betray that trust. However, over time, the tendency of scientists to reproduce and build upon one another's work results in the discovery of the fraudulent data. When that happens, the researchers at fault are usually banished from the community and their careers are ruined.

WATCH **NOW!**

KEY CONCEPT VIDEO 1.8
Solving Chemical Problems

1.8 Solving Chemical Problems

Learning to solve problems is one of the most important skills you will acquire in this course. No one succeeds in chemistry—or in life, really—without the ability to solve problems. Although no simple formula applies to every chemistry problem, you can learn problem-solving strategies and begin to develop some chemical intuition. Many of the problems you will solve in this course are *unit conversion problems*, where you are given one or more quantities and asked to convert them into different units. Other problems require that you use *specific equations* to get to the information you are trying to find. In the sections that follow, you will find strategies to help you solve both of these types of problems. Of course, many problems contain both conversions and equations, requiring the combination of these strategies, and some problems may require an altogether different approach.

Converting from One Unit to Another

In Section 1.6, we discussed the SI unit system, the prefix multipliers, and a few other units. Knowing how to work with and manipulate these units in calculations is central to solving chemical problems. In calculations, units help to determine correctness. Using units as a guide to solving problems is called **dimensional analysis**. Units should always be included in calculations; they are multiplied, divided, and canceled like any other algebraic quantity.

Consider converting 12.5 inches (in) to centimeters (cm). We know from Table 1.3 that 1 in = 2.54 cm (exact), so we can use this quantity in the calculation.

$$12.5\ \cancel{\text{in}} \times \frac{2.54\text{ cm}}{1\ \cancel{\text{in}}} = 31.8\text{ cm}$$

The unit, in, cancels, and we are left with cm as our final unit. The quantity $\frac{2.54\text{ cm}}{1\text{ in}}$ is a **conversion factor**—a fractional quantity with the units we are *converting from* on the bottom and the units we are *converting to* on the top. Conversion factors are constructed from any two equivalent quantities. In this example, 2.54 cm = 1 in, so we construct the conversion factor by dividing both sides of the equality by 1 in and canceling the units.

$$2.54\text{ cm} = 1\text{ in}$$

$$\frac{2.54\text{ cm}}{1\text{ in}} = \frac{1\ \cancel{\text{in}}}{1\ \cancel{\text{in}}}$$

$$\frac{2.54\text{ cm}}{1\text{ in}} = 1$$

Because the quantity $\frac{2.54\text{ cm}}{1\text{ in}}$ is equivalent to 1, multiplying by the conversion factor affects only the units, not the actual quantity. To convert the other way, from centimeters to inches, we must—using units as a guide—use a different form of the conversion factor. If we accidentally use the same form, we will get the wrong result, indicated by erroneous units. For example, suppose that we want to convert 31.8 cm to inches.

$$31.8\text{ cm} \times \frac{2.54\text{ cm}}{1\text{ in}} = \frac{80.8\text{ cm}^2}{\text{in}}$$

The units in the answer (cm^2/in), as well as the value of the answer, are obviously wrong. When we solve a problem, we always look at the final units. Are they the desired units? We must always look at the magnitude of the numerical answer as well. Does it make sense? In this case, our mistake was the form of the conversion factor. It should have been inverted so that the units cancel as follows:

$$31.8\ \cancel{\text{cm}} \times \frac{1\text{ in}}{2.54\ \cancel{\text{cm}}} = 12.5\text{ in}$$

We can invert conversion factors because they are equal to 1 and the inverse of 1 is 1. Therefore,

$$\frac{2.54\text{ cm}}{1\text{ in}} = 1 = \frac{1\text{ in}}{2.54\text{ cm}}$$

Most unit conversion problems take the following form:

$$\text{Information given} \times \text{conversion factor(s)} = \text{information sought}$$

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

In this book, we diagram problem solutions using a *conceptual plan*. A conceptual plan is a visual outline that helps us to see the general flow of the problem solution. For unit conversions, the conceptual plan focuses on units and the conversion from one unit to another. The conceptual plan for converting in to cm is:

$$\boxed{\text{in}} \xrightarrow[\frac{2.54\text{ cm}}{1\text{ in}}]{} \boxed{\text{cm}}$$

The conceptual plan for converting the other way, from cm to in, is just the reverse, with the reciprocal conversion factor:

$$\boxed{\text{cm}} \xrightarrow[\frac{1\text{ in}}{2.54\text{ cm}}]{} \boxed{\text{in}}$$

Each arrow in a conceptual plan for a unit conversion has an associated conversion factor, with the units of the previous step in the denominator and the units of the following step in the numerator. In the following section, we incorporate the idea of a conceptual plan into an overall approach to solving numerical chemical problems.

General Problem-Solving Strategy

In this book, we introduce a standard problem-solving procedure that you can adapt to many of the problems encountered in general chemistry and beyond. One of the difficulties beginning students encounter when trying to solve problems in general chemistry is not knowing where to begin. To solve any problem, you must assess the information given in the problem and devise a way to determine the requested information. In other words, you must do the following:

- Identify the starting point (the ***given*** information).
- Identify the end point (what you must ***find***).
- Devise a way to get from the starting point to the end point using what is given as well as what you already know or can look up. (As we just discussed, this is the ***conceptual plan***.)

In graphic form, we can represent this progression as:

Given $\longrightarrow$ **Conceptual Plan** $\longrightarrow$ **Find**

Although no problem-solving procedure is applicable to all problems, the following four-step procedure can be helpful in working through many of the numerical problems you will encounter in this book.

1. **Sort.** Begin by sorting the information in the problem. *Given* information is the basic data provided by the problem—often one or more numbers with their associated units. *Find* indicates what information you will need for your answer.
2. **Strategize.** This is usually the most challenging part of solving a problem. In this process, you must develop a *conceptual plan*—a series of steps that take you from the given information to the information you are trying to find. You have already seen conceptual plans for simple unit conversion problems. Each arrow in a conceptual plan represents a computational step. On the left side of the arrow is the quantity you had before the step, on the right side of the arrow is the quantity you have after the step, and below the arrow is the information you need to get from one to the other—the relationship between the quantities.

Most problems can be solved in more than one way. The solutions in this book tend to be the most straightforward but certainly are not the only way to solve the problem.

Often such relationships take the form of conversion factors or equations. These may be given in the problem, in which case you will have written them down under "Given" in step 1. Usually, however, you will need other information—which may include physical constants, formulas, or conversion factors—to help get you from what you are given to what you must find. This information comes from what you have learned or can look up in the chapter or in tables within the book.

In some cases, you may get stuck at the strategize step. If you cannot figure out how to get from the given information to the information you are asked to find, you might try working backward. For example, you can look at the units of the quantity you are trying to find and attempt to find conversion factors to get to the units of the given quantity. You may even try a combination of strategies; work forward, backward, or some of both. If you persist, you will develop a strategy to solve the problem.

3. **Solve.** This is the easiest part of solving a problem. Once you set up the problem properly and devise a conceptual plan, you simply follow the plan to solve the problem. Carry out any mathematical operations (paying attention to the rules for significant figures in calculations) and cancel units as needed.
4. **Check.** This is the step beginning students most often overlook. Experienced problem solvers always ask, "Does this answer make sense? Are the units correct? Is the number of significant figures correct?" When solving multistep problems, errors can easily creep into the solution. You can catch most of these errors by simply

checking the answer. For example, suppose you are calculating the number of atoms in a gold coin and end up with an answer of 1.1×10^{-6} atoms. Could the gold coin really be composed of one-millionth of one atom?

In Examples 1.7 and 1.8, this problem-solving procedure is applied to unit conversion problems. The procedure is summarized in the left column, and two examples of applying the procedure are provided in the middle and right columns. This three-column format is used in selected examples throughout this text. It allows you to see how to apply a particular procedure to two different problems. Work through one problem first (from top to bottom) and then see how you can apply the same procedure to the other problem. Recognizing the commonalities and differences between problems is a key part of developing problem-solving skills.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE VIDEO 1.8

HOW TO: *Solve Unit Conversion Problems*	**Unit Conversion** **EXAMPLE 1.7** Convert 1.76 yards to centimeters.	**Unit Conversion** **EXAMPLE 1.8** Convert 1.8 quarts to cubic centimeters.
SORT Begin by sorting the information in the problem into *given* and *find*.	**GIVEN:** 1.76 yd **FIND:** cm	**GIVEN:** 1.8 qt **FIND:** cm^3
STRATEGIZE Devise a *conceptual plan* for the problem. Begin with the *given* quantity and symbolize each conversion step with an arrow. Below each arrow, write the appropriate conversion factor for that step. Focus on the units. The conceptual plan should end at the *find* quantity and its units. In these examples, the other information you need consists of relationships between the various units.	**CONCEPTUAL PLAN** $\text{yd} \xrightarrow[\frac{1\text{ m}}{1.094\text{ yd}}]{} \text{m} \xrightarrow[\frac{1\text{ cm}}{10^{-2}\text{ m}}]{} \text{cm}$ **RELATIONSHIPS USED** $1.094\text{ yd} = 1\text{ m}$ $1\text{ cm} = 10^{-2}\text{ m}$ (These conversion factors are from Tables 1.2 and 1.3.)	**CONCEPTUAL PLAN** $\text{qt} \xrightarrow[\frac{1\text{ L}}{1.057\text{ qt}}]{} \text{L} \xrightarrow[\frac{1\text{ mL}}{10^{-3}\text{ L}}]{} \text{mL} \xrightarrow[\frac{1\text{ cm}^3}{1\text{ mL}}]{} \text{cm}^3$ **RELATIONSHIPS USED** $1.057\text{ qt} = 1\text{ L}$ $1\text{ mL} = 10^{-3}\text{ L}$ $1\text{ mL} = 1\text{ cm}^3$ (These conversion factors are from Tables 1.2 and 1.3.)
SOLVE Follow the conceptual plan. Begin with the *given* quantity and its units. Multiply by the appropriate conversion factor(s), canceling units, to arrive at the *find* quantity. Round the answer to the correct number of significant figures following the rules in Section 1.7. Remember that exact conversion factors do not limit significant figures.	**SOLUTION** $1.76\ \cancel{\text{yd}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1\text{ cm}}{10^{-2}\ \cancel{\text{m}}} = 160.8775\text{ cm}$ $160.8775\text{ cm} = 161\text{ cm}$	**SOLUTION** $1.8\ \cancel{\text{qt}} \times \frac{1\ \cancel{\text{L}}}{1.057\ \cancel{\text{qt}}} \times \frac{1\ \cancel{\text{mL}}}{10^{-3}\ \cancel{\text{L}}} \times \frac{1\text{ cm}^3}{1\ \cancel{\text{mL}}} = 1.70293 \times 10^3\text{ cm}^3$ $1.70293 \times 10^3\text{ cm}^3 = 1.7 \times 10^3\text{ cm}^3$
CHECK Check your answer. Are the units correct? Does the answer make sense?	The units (cm) are correct. The magnitude of the answer (161) makes sense because a centimeter is a much smaller unit than a yard.	The units (cm^3) are correct. The magnitude of the answer (1700) makes sense because a cubic centimeter is a much smaller unit than a quart.
	FOR PRACTICE 1.7 Convert 288 cm to yards.	**FOR PRACTICE 1.8** Convert 9255 cm^3 to gallons.

Units Raised to a Power

When building conversion factors for units raised to a power, remember to raise both the number and the unit to the power. For example, to convert from in^2 to cm^2, we construct the conversion factor as follows:

$$\begin{aligned} 2.54\ \text{cm} &= 1\ \text{in} \\ (2.54\ \text{cm})^2 &= (1\ \text{in})^2 \\ (2.54)^2\ \text{cm}^2 &= 1^2\ \text{in}^2 \\ 6.45\ \text{cm}^2 &= 1\ \text{in}^2 \\ \frac{6.45\ \text{cm}^2}{1\ \text{in}^2} &= 1 \end{aligned}$$

Example 1.9 demonstrates how to use conversion factors involving units raised to a power.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 1.9

EXAMPLE 1.9 Unit Conversions Involving Units Raised to a Power

Calculate the displacement (the total volume of the cylinders through which the pistons move) of a 5.70-L automobile engine in cubic inches.

SORT Sort the information in the problem into *given* and *find*.

GIVEN: 5.70 L
FIND: in^3

STRATEGIZE Write a conceptual plan. Begin with the given information and devise a path to the information that you are asked to find. For cubic units, you must cube the conversion factors.

CONCEPTUAL PLAN

$$\text{L} \xrightarrow{\frac{1\ \text{mL}}{10^{-3}\ \text{L}}} \text{mL} \xrightarrow{\frac{1\ \text{cm}^3}{1\ \text{mL}}} \text{cm}^3 \xrightarrow{\frac{(1\ \text{in})^3}{(2.54\ \text{cm})^3}} \text{in}^3$$

RELATIONSHIPS USED

$$\begin{aligned} 1\ \text{mL} &= 10^{-3}\ \text{L} \\ 1\ \text{mL} &= 1\ \text{cm}^3 \\ 2.54\ \text{cm} &= 1\ \text{in} \end{aligned}$$

(These conversion factors are from Tables 1.2 and 1.3.)

SOLVE Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the least precisely known quantity (5.70 L). These conversion factors are all exact and therefore do not limit the number of significant figures.

SOLUTION

$$5.70\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mL}}}{10^{-3}\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \frac{(1\ \text{in})^3}{(2.54\ \cancel{\text{cm}})^3}$$
$$= 347.835\ \text{in}^3 = 348\ \text{in}^3$$

CHECK The units of the answer are correct, and the magnitude makes sense. The unit cubic inches is smaller than liters, so the volume in cubic inches should be larger than the volume in liters.

FOR PRACTICE 1.9 How many cubic centimeters are there in 2.11 yd^3?

FOR MORE PRACTICE 1.9 A vineyard has 145 acres of Chardonnay grapes. A particular soil supplement requires 5.50 g for every square meter of vineyard. How many kilograms of the soil supplement are required for the entire vineyard? (1 km^2 = 247 acres)

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE 1.10

EXAMPLE 1.10 Density as a Conversion Factor

The mass of fuel in a jet must be calculated before each flight to ensure that the jet is not too heavy to fly. A 747 is fueled with 173,231 L of jet fuel. If the density of the fuel is 0.768 g/cm^3, what is the mass of the fuel in kilograms?

SORT Begin by sorting the information in the problem into *given* and *find*.

GIVEN: fuel volume = 173,231 L
density of fuel = 0.768 g/cm^3
FIND: mass in kg

STRATEGIZE Draw the conceptual plan by beginning with the given quantity—in this case the volume in liters (L). The overall goal of this problem is to find the mass. You can convert between volume and mass using density (g/cm^3). However, you must first convert the volume to cm^3. After you convert the volume to cm^3, use the density to convert to g. Finally, convert g to kg.

CONCEPTUAL PLAN

$$\text{L} \xrightarrow{\frac{1\text{ mL}}{10^{-3}\text{ L}}} \text{mL} \xrightarrow{\frac{1\text{ cm}^3}{1\text{ mL}}} \text{cm}^3 \xrightarrow{\frac{0.768\text{ g}}{\text{cm}^3}} \text{g} \xrightarrow{\frac{1\text{ kg}}{1000\text{ g}}} \text{kg}$$

RELATIONSHIPS USED

$1\text{ mL} = 10^{-3}\text{ L}$

$1\text{ mL} = 1\text{ cm}^3$

$d = 0.768\text{ g/cm}^3$

$1000\text{ g} = 1\text{ kg}$

(These conversion factors are from Tables 1.2 and 1.3.)

SOLVE Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the density.

SOLUTION

$$173{,}231\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mL}}}{10^{-3}\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \frac{0.768\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}} \times \frac{1\text{ kg}}{1000\ \cancel{\text{g}}}$$

$$= 1.33 \times 10^5\text{ kg}$$

CHECK The units of the answer (kg) are correct. The magnitude makes sense because the mass (1.33×10^5 kg) is similar in magnitude to the given volume (173,231 L or 1.73231×10^5 L), as expected for a density close to 1 (0.768 g/cm^3).

FOR PRACTICE 1.10 Backpackers often use canisters of white gas to fuel a cooking stove's burner. If one canister contains 1.45 L of white gas and the density of the gas is 0.710 g/cm^3, what is the mass of the fuel in kilograms?

FOR MORE PRACTICE 1.10 A drop of gasoline has a mass of 22 mg and a density of 0.754 g/cm^3. What is its volume in cubic centimeters?

Order-of-Magnitude Estimations

Calculation is an integral part of chemical problem solving. But precise numerical calculation is not always necessary, or even possible. Sometimes data are only approximate; other times we may not need a high degree of precision—a rough estimate or a simplified "back of the envelope" calculation is enough. We can also use approximate calculations to get an initial feel for a problem or to make a quick check to see whether our solution is "in the right ballpark."

One way to make such estimates is to simplify the numbers so that they can be manipulated easily. The technique known as *order-of-magnitude estimation* is based on focusing only on the exponential part of numbers written in scientific notation, according to these guidelines:

- If the decimal part of the number is less than 5, just drop it. Thus, 4.36×10^5 becomes 10^5 and 2.7×10^{-3} becomes 10^{-3}.
- If the decimal part is 5 or more, round it up to 10 and rewrite the number as a power of 10. Thus, 5.982×10^7 becomes $10 \times 10^7 = 10^8$, and 6.1101×10^{-3} becomes $10 \times 10^{-3} = 10^{-2}$.

When we make these approximations, we are left with powers of 10, which are easier to multiply and divide. Of course, our answer is only as reliable as the numbers used to get it, so we should not assume that the results of an order-of-magnitude calculation are accurate to more than an order of magnitude.

Suppose, for example, that we want to estimate the number of atoms that an immortal being could have counted in the 14 billion (1.4×10^{10}) years that the universe has been in existence, assuming a counting rate of ten atoms per second. Since a year has 3.2×10^7 seconds, we can approximate the number of atoms counted as follows:

$$\underbrace{10^{10}\ \cancel{\text{years}}}_{\text{(number of years)}} \times \underbrace{10^7\ \frac{\cancel{\text{seconds}}}{\cancel{\text{years}}}}_{\text{(number of seconds per year)}} \times \underbrace{10^1\ \frac{\text{atoms}}{\cancel{\text{second}}}}_{\text{(number of atoms counted per second)}} \approx 10^{18}\text{ atoms}$$

A million trillion atoms (10^{18}) may seem like a lot, but as we discuss in Chapter 2, a speck of matter made up of a million trillion atoms is nearly impossible to see without a microscope.

In our general problem-solving procedure, the last step is to check whether the results seem reasonable. Order-of-magnitude estimations often help us catch the kinds of mistakes that can happen in a detailed calculation, such as entering an incorrect exponent or sign into a calculator or multiplying when we should have divided.

Problems Involving an Equation

We can solve problems involving equations in much the same way as problems involving conversions. Usually, in problems involving equations, we must find one of the variables in the equation, given the others. The *conceptual plan* concept outlined earlier can be used for problems involving equations. For example, suppose we are given the mass (m) and volume (V) of a sample and asked to calculate its density. The conceptual plan shows how the *equation* takes us from the *given* quantities to the *find* quantity:

$$m, V \xrightarrow[d = \frac{m}{V}]{} d$$

Here, instead of a conversion factor under the arrow, this conceptual plan has an equation. The equation shows the *relationship* between the quantities on the left of the arrow and the quantities on the right. Note that at this point the equation need not be solved for the quantity on the right (although in this particular case it is). The procedure that follows, as well as the two examples, offer guidance for developing a strategy to solve problems involving equations. We again use the three-column format. Work through one problem from top to bottom and then see how you can apply the same general procedure to the second problem.

WATCH **NOW!**
INTERACTIVE WORKED EXAMPLE VIDEO 1.12

HOW TO: *Solve Problems Involving Equations*	**Problems with Equations**	**Problems with Equations**
	EXAMPLE 1.11 Find the radius (r) in centimeters of a spherical water droplet with a volume (V) of 0.058 cm^3. For a sphere, $V = (4/3)\,\pi r^3$.	**EXAMPLE 1.12** Find the density (in g/cm^3) of a metal cylinder with a mass (m) of 8.3 g, a length (l) of 1.94 cm, and a radius (r) of 0.55 cm. For a cylinder, $V = \pi r^2 l$.
SORT Begin by sorting the information in the problem into *given* and *find*.	**GIVEN:** $V = 0.058\ cm^3$ **FIND:** r in cm	**GIVEN:** $m = 8.3$ g $l = 1.94$ cm $r = 0.55$ cm **FIND:** d in g/cm^3
STRATEGIZE Write a conceptual plan for the problem. Focus on the equation(s). The conceptual plan shows how the equation takes you from the *given* quantity (or quantities) to the *find* quantity. The conceptual plan may have several parts, involving other equations or required conversions. In these examples, you use the geometrical relationships given in the problem statements as well as the definition of density, $d = m/V$, which you learned in this chapter.	**CONCEPTUAL PLAN** $V \xrightarrow[V = \frac{4}{3}\pi r^3]{} r$ **RELATIONSHIPS USED** $V = \frac{4}{3}\pi r^3$	**CONCEPTUAL PLAN** $l, r \xrightarrow[V = \pi r^2 l]{} V$, $m, V \xrightarrow[d = m/V]{} d$ **RELATIONSHIPS USED** $V = \pi r^2 l$ $d = \frac{m}{V}$

SOLVE Follow the conceptual plan. Solve the equation(s) for the *find* quantity (if it is not already). Gather each of the quantities that must go into the equation in the correct units. (Convert to the correct units if necessary.) Substitute the numerical values and their units into the equation(s) and calculate the answer. Round the answer to the correct number of significant figures.	**SOLUTION** $V = \frac{4}{3}\pi r^3$ $r^3 = \frac{3}{4\pi}V$ $r = \left(\frac{3}{4\pi}V\right)^{1/3}$ $= \left(\frac{3}{4\pi}0.058\text{ cm}^3\right)^{1/3}$ $= 0.24013\text{ cm}$ $0.24013\text{ cm} = 0.24\text{ cm}$	**SOLUTION** $V = \pi r^2 l$ $= \pi(0.55\text{ cm})^2(1.94\text{ cm})$ $= 1.8\underline{4}36\text{ cm}^3$ $d = \frac{m}{V}$ $= \frac{8.3\text{ g}}{1.8\underline{4}36\text{ cm}^3} = 4.50195\text{ g/cm}^3$ $4.50195\text{ g/cm}^3 = 4.5\text{ g/cm}^3$
CHECK Check your answer. Are the units correct? Does the answer make sense?	The units (cm) are correct and the magnitude makes sense.	The units (g/cm^3) are correct. The magnitude of the answer seems correct for one of the lighter metals (see Table 1.4).
	FOR PRACTICE 1.11 Find the radius (r) of an aluminum cylinder that is 2.00 cm long and has a mass of 12.4 g. For a cylinder, $V = \pi r^2 l$.	**FOR PRACTICE 1.12** Find the density, in g/cm^3, of a metal cube with a mass of 50.3 g and an edge length (l) of 2.65 cm. For a cube, $V = l^3$.

1.9 Analyzing and Interpreting Data

In Section 1.2, we saw that Lavoisier saw patterns in a series of related measurements. Sets of measurements constitute scientific *data*, and learning to analyze and interpret data is an important scientific skill.

Identifying Patterns in Data

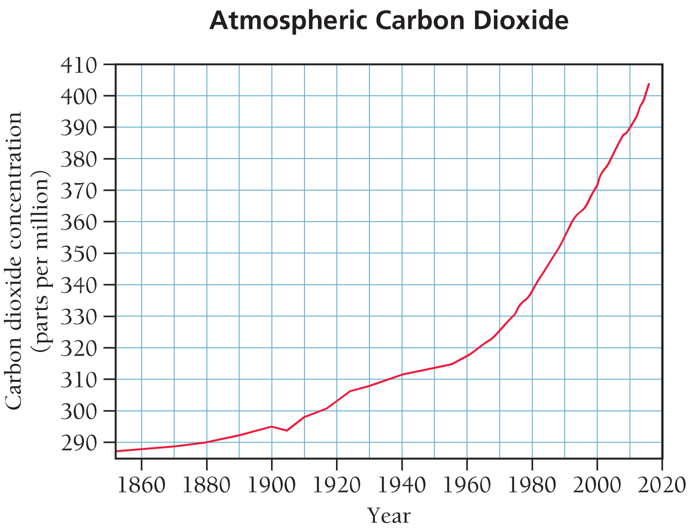
Suppose you are an early chemist trying to understand the composition of water. You know that water is composed of the elements hydrogen and oxygen. You do several experiments in which different samples of water are decomposed into hydrogen and oxygen and you get the following results:

Sample	Mass of Water Sample	Mass of Hydrogen Formed	Mass of Oxygen Formed
A	20.0 g	2.2 g	17.8 g
B	50.0 g	5.6 g	44.4 g
C	100.0 g	11.1 g	88.9 g

Do you notice any patterns in these data? Perhaps the easiest pattern to see is that the sum of the masses of oxygen and hydrogen always sum to the mass of the water sample. For example, for the first water sample, 2.2 g hydrogen + 17.8 g oxygen = 20.0 g water. The pattern is the same for the other samples. Another pattern, which is a bit more difficult to see, is that the ratio of the masses of oxygen and hydrogen is the same for each sample.

Sample	Mass of Hydrogen Formed	Mass of Oxygen Formed	$\frac{\text{Mass Oxygen}}{\text{Mass Hydrogen}}$
A	2.2 g	17.8 g	8.1
B	5.6 g	44.4 g	7.9
C	11.1 g	88.9 g	8.01

The ratio is 8—the small variations are due to experimental error, which is common in all measurements and observations.



▲ **FIGURE 1.14 Atmospheric Carbon Dioxide Levels from 1860 to Present.**

Seeing patterns in data is a creative process that requires you to not just merely tabulate laboratory measurements, but to see relationships that may not always be obvious. The best scientists see patterns that others have missed. As you learn to interpret data in this course, be creative and try looking at data in new ways.

Interpreting Graphs

Data are often visualized using graphs or images and scientists must constantly analyze and interpret graphs. For example, the graph in Figure 1.14◀ shows the concentration of carbon dioxide in Earth's atmosphere as a function of time. Carbon dioxide is a greenhouse gas that has been rising as result of the burning of fossil fuels (such as gasoline and coal). When you look at a graph such as this one, you should first examine the *x* and *y* axes and make sure you understand what each axis represents. You should also examine the numerical range of the axes. In Figure 1.14 the *y* axis does not begin at zero in order to better display the change that is occurring. How would this graph look different if the *y* axis began at zero instead of at 290? Notice also that, in this graph, the increase in carbon dioxide has not been constant over time. The rate of increase—represented by the slope of the line—has intensified since about 1960.

EXAMPLE 1.13 Interpreting Graphs

Examine the graph in Figure 1.14 and answer each question.

(a) What was the concentration of carbon dioxide in 1960?

(b) What was the concentration in 2010?

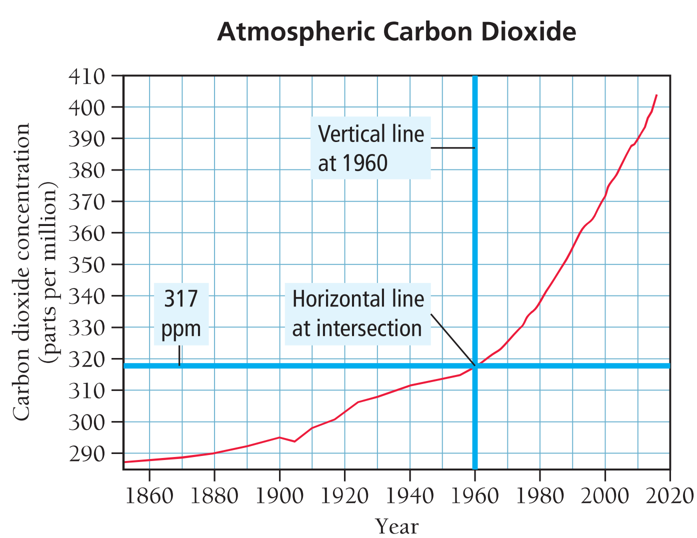
(c) How much did the concentration increase between 1960 and 2010?

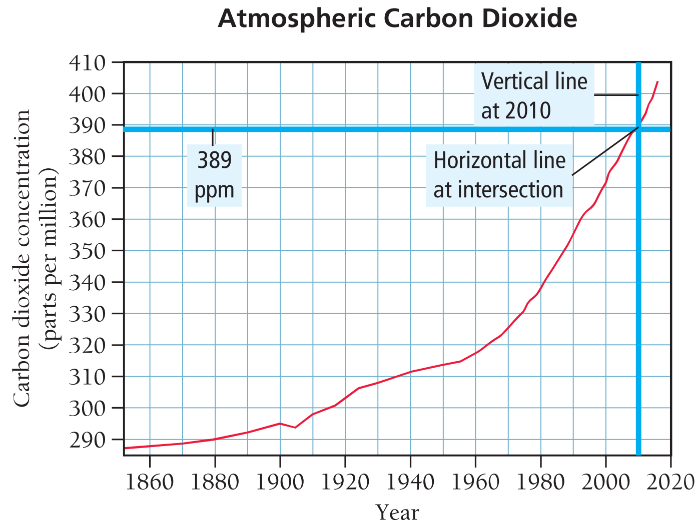
(d) What is the average rate of increase over this time?

(e) If the average rate of increase from part d remains constant, estimate the carbon dioxide concentration in 2050. (Use the concentration in 2010 as your starting point.)

SOLUTION

(a) To determine the concentration of carbon dioxide in 1960, draw a vertical line at the year 1960. At the point where the vertical line intersects the carbon dioxide concentration curve, draw a horizontal line. The point where the horizontal line intercepts the *y* axis represents the concentration in 1960. So, the concentration in 1960 was 317 ppm.



(b) Apply the same procedure as in part a, but now shift the vertical line to the year 2010. The concentration in the year 2010 was 389 ppm.	
(c) The increase in the carbon dioxide concentration is the difference between the two concentrations. When calculating changes in quantities such as this, take the final quantity minus the initial quantity.	Increase in concentration = concentration in 2010 − concentration in 1960 = 389 ppm − 317 ppm = 72 ppm
(d) The average rate of increase over this time is the change in the concentration divided by the number of years that passed. Determine the number of years that have passed, which is equal to the final year minus the initial year.	Number of years = final year − initial year = 2010 − 1960 = 50 years
Determine the average rate of increase by dividing the change in concentration from part c by the number of years that you just calculated.	$\text{Average rate} = \dfrac{\text{change in concentration}}{\text{number of years}} = \dfrac{72\text{ ppm}}{50\text{ years}} = \dfrac{1.4\text{ ppm}}{\text{year}}$
(e) Determine the increase in concentration between 2010 and 2050 by multiplying the number of years that pass in that time interval by the average rate of change from part d. Lastly, determine the concentration in 2050 by adding the increase between 2010 and 2050 to the concentration in 2010.	$\text{Increase} = 40\text{ years} \times \dfrac{1.4\text{ ppm}}{\text{year}} = 56\text{ ppm}$ Concentration in 2050 = 389 ppm + 56 ppm = 445 ppm

FOR PRACTICE 1.13 What was the average rate of increase in carbon dioxide concentration between 1880 and 1920? Why might that rate be different than the rate between 1960 and 2010?

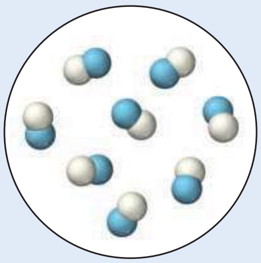
QUIZ YOURSELF **NOW!**

Self-Assessment Quiz

Q1. A chemist mixes sodium with water and witnesses a violent reaction between the metal and water. This is best classified as
MISSED THIS? *Read Section 1.2*

a) an observation. b) a law.
c) a hypothesis. d) a theory.

Q2. This image represents a particulate view of a sample of matter. Classify the sample according to its composition.
MISSED THIS? *Read Section 1.3; Watch KCV 1.3*



a) The sample is a pure element.
b) The sample is a homogeneous mixture.
c) The sample is a compound.
d) The sample is a heterogeneous mixture.

Q3. Which change is a physical change?
MISSED THIS? *Read Section 1.4*

a) wood burning b) iron rusting
c) dynamite exploding d) gasoline evaporating

Q4. Which property of rubbing alcohol is a chemical property?
MISSED THIS? *Read Section 1.4*

a) density (0.786 g/cm^3) b) flammability
c) boiling point (82.5 °C) d) melting point (–89 °C)

Q5. Convert 85.0 °F to K.
MISSED THIS? *Read Section 1.6; Watch KCV 1.3*

a) 181.1 K b) 358 K
c) 29.4 K d) 302.6 K

Q6. Express the quantity 33.2×10^{-4} m in mm.
MISSED THIS? *Read Section 1.6; Watch KCV 1.3*

a) 33.2 mm b) 3.32 mm
c) 0.332 mm d) 3.32×10^{-6} mm

Q7. What is the mass of a 1.75 L sample of a liquid that has a density of 0.921 g/mL?
MISSED THIS? *Read Section 1.6; Watch KCV 1.10*

a) 1.61×10^3 g b) 1.61×10^{-3} g
c) 1.90×10^3 g d) 1.90×10^{-3} g

Q8. Perform the calculation to the correct number of significant figures. **MISSED THIS?** *Read Section 1.7; Watch KCV 1.6, 1.7, IWE 1.5, 1.6*

$$(43.998 \times 0.00552)/2.002$$

a) 0.121 b) 0.12
c) 0.12131 d) 0.1213

Q9. Perform the calculation to the correct number of significant figures. **MISSED THIS?** *Read Section 1.7; Watch KCV 1.6, 1.7, IWE 1.5, 1.6*

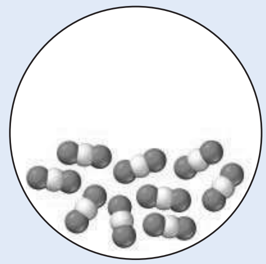
$$(8.01 - 7.50)/3.002$$

a) 0.1698867 b) 0.17
c) 0.170 d) 0.1699

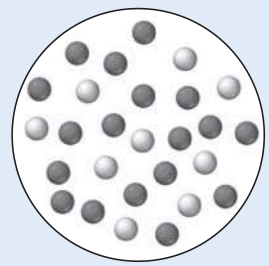
Q10. Convert 1285 cm^2 to m^2.
MISSED THIS? *Read Section 1.8; Watch IWE 1.9*

a) 1.285×10^7 m^2 b) 12.85 m^2
c) 0.1285 m^2 d) 1.285×10^5 m^2

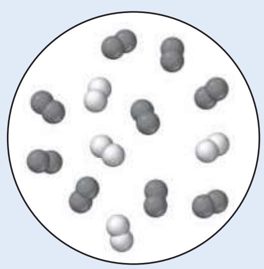
Q11. The first diagram depicts a compound in its liquid state. Which of the other diagrams best depicts the compound after it has evaporated into a gas?
MISSED THIS? *Read Section 1.4*



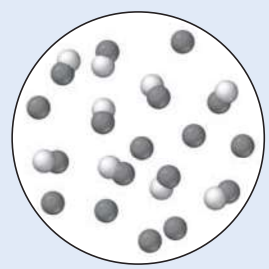
(a)



(b)



(c)



(d)

Q12. Three samples, each of a different substance, are weighed, and their volume is measured. The results are tabulated. List the substances in order of decreasing density.
MISSED THIS? *Read Section 1.6*

	Mass	Volume
Substance I	10.0 g	10.0 mL
Substance II	10.0 kg	12.0 L
Substance III	12.0 mg	10.0 μL

a) III > II > I
b) I > II > III
c) III > I > II
d) II > I > III

Q13. A solid metal sphere has a radius of 3.53 cm and a mass of 1.796 kg. What is the density of the metal in g/cm^3? (The volume of a sphere is $V = \frac{4}{3}\pi r^3$.)
MISSED THIS? *Read Section 1.6; Watch IWE 1.12*

a) $34.4\ g/cm^3$ b) $0.103\ g/cm^3$
c) $121\ g/cm^3$ d) $9.75\ g/cm^3$

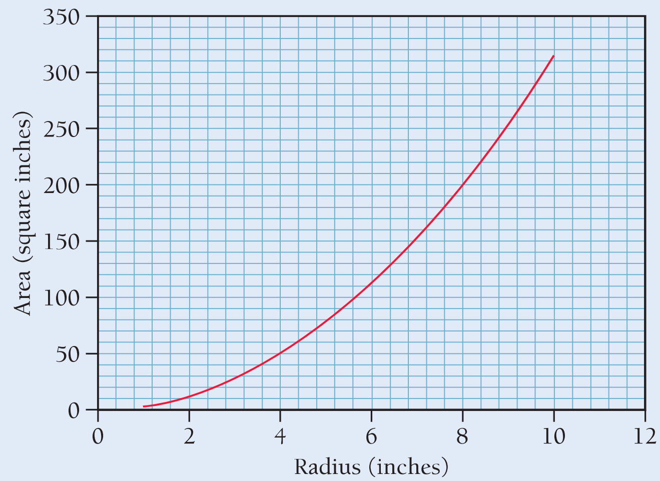
Q14. The gas mileage of a certain German automobile is 22 km/L. Convert this quantity to miles per gallon.
MISSED THIS? *Read Section 1.8; Watch KCV 1.8*

a) 9.4 mi/gal b) 1.3×10^2 mi/gal
c) 52 mi/gal d) 3.6 mi/gal

Q15. A wooden block has a volume of $18.5\ in^3$. Express the volume of the cube in cm^3.
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.9*

a) $303\ cm^3$ b) $47.0\ cm^3$
c) $1.13\ cm^3$ d) $7.28\ cm^3$

Q16. The graph that follows shows the area of a circle as a function of its radius. What is the radius of a circle that has an area of 155 square inches?
MISSED THIS? *Read Section 1.9*



a) 7.0 inches b) 6.5 inches
c) 6.8 inches d) 6.2 inches

Answers: **1.** (a) **2.** (c) **3.** (d) **4.** (b) **5.** (d) **6.** (b) **7.** (a) **8.** (a) **9.** (b) **10.** (c) **11.** (a) **12.** (c) **13.** (d) **14.** (c) **15.** (a) **16.** (a)

CHAPTER 1 IN REVIEW

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CONCEPTS

Atoms and Molecules (1.1)

- All matter is composed of atoms and molecules.
- Chemistry is the science that investigates the properties of matter by examining the atoms and molecules that compose it.

The Scientific Approach to Knowledge (1.2)

- Science begins with the observation of the physical world. A number of related observations can be summarized in a statement or generalization called a scientific law.
- A hypothesis is a tentative interpretation or an explanation of observations. One or more well-established hypotheses may prompt the development of a scientific theory, a model for nature that explains the underlying reasons for observations and laws.
- Laws, hypotheses, and theories all give rise to predictions that can be tested by experiments, carefully controlled procedures designed to produce critical new observations. If scientists cannot confirm the predictions, they must modify or replace the law, hypothesis, or theory.

The Classification of Matter (1.3)

- We classify matter according to its state (solid, liquid, or gas) or according to its composition (pure substance or mixture).
- A pure substance can either be an element, which cannot be chemically broken down into simpler substances, or a compound, which is composed of two or more elements in fixed proportions.
- A mixture can be either homogeneous, with the same composition throughout, or heterogeneous, with different compositions in different regions.

The Properties of Matter (1.4)

- We classify the properties of matter into two types: physical and chemical. Matter displays its physical properties without changing its composition.
- Changes in matter in which composition does not change are physical changes. Changes in matter in which composition does change are chemical changes.

Energy (1.5)

- In chemical and physical changes, matter often exchanges energy with its surroundings. In these exchanges, the total energy is always conserved; energy is neither created nor destroyed.
- Systems with high potential energy tend to change in the direction of lower potential energy, releasing energy into the surroundings.

The Units of Measurement and Significant Figures (1.6, 1.7)

- Scientists use SI units, which are based on the metric system. The SI base units include the meter (m) for length, the kilogram (kg) for mass, the second (s) for time, and the kelvin (K) for temperature.
- Derived units are formed from a combination of other units. Common derived units include those for volume (cm^3 or m^3) and density (g/cm^3).
- The number of digits in a reported measurement reflects the uncertainty in the measurement. Significant figures are the non–place-holding digits in a reported number.

EQUATIONS AND RELATIONSHIPS

Relationship between Kelvin (K) and Celsius (°C) Temperature Scales (1.6)

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

Relationship between Celsius (°C) and Fahrenheit (°F) Temperature Scales (1.6)

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

Relationship between Density (d), Mass (m), and Volume (V) (1.6)

$$d = \frac{m}{V}$$

LEARNING OUTCOMES

Chapter Objectives	Assessment
Apply the scientific approach (1.2)	Exercises 33–36
Classify matter according to its composition (1.3)	Exercises 37–42
Classify the properties and changes in matter as chemical or physical (1.4)	Example 1.1 **For Practice** 1.1 **Exercises** 43–50
Compare the Fahrenheit, Celsius, and Kelvin temperature scales (1.6)	Example 1.2 **For Practice** 1.2 **Exercises** 51–54
Express measurements using appropriate prefix multipliers (1.6)	Exercises 55–64
Apply the density relationship to problems involving mass and volume (1.6)	Example 1.3 **For Practice** 1.3 **For More Practice** 1.3 **Exercises** 65–72
Determine the number of significant figures in a measurement or reported number (1.7)	Examples 1.4, 1.5 **For Practice** 1.4, 1.5 **Exercises** 73–82
Determine the number of significant figures in the result of a given calculation (1.7)	Example 1.6 **For Practice** 1.6 **Exercises** 83–90
Convert between units using dimensional analysis (1.8)	Examples 1.7, 1.8, 1.9, 1.10 **For Practice** 1.7, 1.8, 1.9, 1.10 **For More Practice** 1.9, 1.10 **Exercises** 91–104
Solve problems involving equations (1.8)	Examples 1.11, 1.12 **For Practice** 1.11, 1.12

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. Explain this statement in your own words and give an example. *The properties of the substances around us depend on the atoms and molecules that compose them.*

2. Explain the main goal of chemistry.

3. Describe the scientific approach to knowledge. How does it differ from other approaches?

4. Explain the differences between a hypothesis, a law, and a theory.

5. What observations did Antoine Lavoisier make? What law did he formulate?

6. What theory did John Dalton formulate?

7. What is wrong with the expression "That is just a theory," if by *theory* the speaker is referring to a scientific theory?

8. What are two different ways to classify matter?

9. How do solids, liquids, and gases differ?

10. What is the difference between a crystalline solid and an amorphous solid?

11. Explain the difference between a pure substance and a mixture.

12. Explain the difference between an element and a compound.

13. Explain the difference between a homogeneous and a heterogeneous mixture.

14. What kind of mixtures can be separated by filtration?

15. Explain how distillation is used to separate mixtures.

16. What is the difference between a physical property and a chemical property?

17. What is the difference between a physical change and a chemical change? List some examples of each.

18. Explain the significance of the law of conservation of energy.

19. What kind of energy is chemical energy? In what way is an elevated weight similar to a tank of gasoline?

20. What are the standard SI base units of length, mass, time, and temperature?

21. What are the three common temperature scales? Does the size of a degree differ among them?

22. What are prefix multipliers? List some examples.

23. What is a derived unit? List an example.

24. Explain the difference between density and mass.

25. Explain the difference between *intensive* and *extensive* properties.

26. What is the meaning of the number of digits reported in a measured quantity?

27. When multiplying or dividing measured quantities, what determines the number of significant figures in the result?

28. When adding or subtracting measured quantities, what determines the number of significant figures in the result?

29. What are the rules for rounding off the results of calculations?

30. Explain the difference between precision and accuracy.

31. Explain the difference between random error and systematic error.

32. What is dimensional analysis?

PROBLEMS BY TOPIC

Note: *Answers to all odd-numbered Problems, numbered in blue, can be found in Appendix III. Exercises in the Problems by Topic section are paired, with each odd-numbered problem followed by a similar even-numbered problem. Exercises in the Cumulative Problems section are also paired, but more loosely. Challenge Problems and Conceptual Problems, because of their nature, are unpaired.*

The Scientific Approach to Knowledge

33. Classify each statement as an observation, a law, or a theory. **MISSED THIS?** *Read Section 1.2*

a. All matter is made of tiny, indestructible particles called atoms.

b. When iron rusts in a closed container, the mass of the container and its contents does not change.

c. In chemical reactions, matter is neither created nor destroyed.

d. When a match burns, heat is released.

34. Classify each statement as an observation, a law, or a theory.

a. Chlorine is a highly reactive gas.

b. If elements are listed in order of increasing mass of their atoms, their chemical reactivities follow a repeating pattern.

c. Neon is an inert (or nonreactive) gas.

d. The reactivity of elements depends on the arrangement of their electrons.

35. A chemist decomposes several samples of carbon monoxide into carbon and oxygen and weighs the resultant elements. The results are shown in the table. **MISSED THIS?** *Read Section 1.2*

Sample	Mass of Carbon (g)	Mass of Oxygen (g)
1	6	8
2	12	16
3	18	24

a. Do you notice a pattern in these results?

Next, the chemist decomposes several samples of hydrogen peroxide into hydrogen and oxygen. The results are shown in the table.

Sample	Mass of Hydrogen (g)	Mass of Oxygen (g)
1	0.5	8
2	1	16
3	1.5	24

b. Do you notice a similarity between these results and those for carbon monoxide in part a?

c. Can you formulate a law from your observations in a and b?

d. Can you formulate a hypothesis that might explain your law in c?

36. When astronomers observe distant galaxies, they can tell that most of them are moving away from one another. In addition, the more distant the galaxies, the more rapidly they are likely to be moving away from each other. Can you devise a hypothesis to explain these observations?

The Classification and Properties of Matter

37. Classify each substance as a pure substance or a mixture. If it is a pure substance, classify it as an element or a compound. If it is a mixture, classify it as homogeneous or heterogeneous.
MISSED THIS? *Read Section 1.3; Watch KCV 1.3*

a. sweat **b.** carbon dioxide
c. aluminum **d.** vegetable soup

38. Classify each substance as a pure substance or a mixture. If it is a pure substance, classify it as an element or a compound. If it is a mixture, classify it as homogeneous or heterogeneous.

a. wine **b.** beef stew
c. iron **d.** carbon monoxide

39. Complete the table.
MISSED THIS? *Read Section 1.3; Watch KCV 1.3*

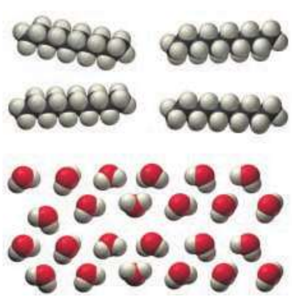
Substance	Pure or mixture	Type
aluminum	pure	element
apple juice	_______	_______
hydrogen peroxide	_______	_______
chicken soup	_______	_______

40. Complete the table.

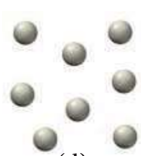
Substance	Pure or mixture	Type
water	pure	compound
coffee	_______	_______
ice	_______	_______
carbon	_______	_______

41. Determine whether each molecular diagram represents a pure substance or a mixture. If it represents a pure substance, classify the substance as an element or a compound. If it represents a mixture, classify the mixture as homogeneous or heterogeneous.
MISSED THIS? *Read Section 1.3; Watch KCV 1.3*

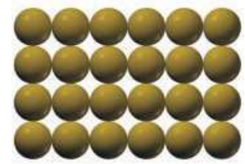

(a)

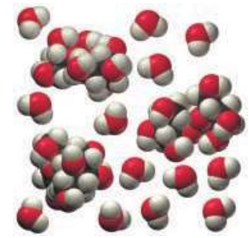

(b)

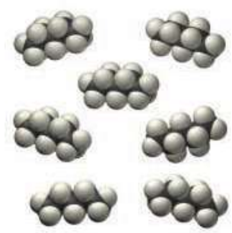

(c)


(d)

42. Determine whether each molecular diagram represents a pure substance or a mixture. If it represents a pure substance, classify the substance as an element or a compound. If it represents a mixture, classify the mixture as homogeneous or heterogeneous.


(a)


(b)


(c)


(d)

43. Classify each of the listed properties of isopropyl alcohol (also known as rubbing alcohol) as physical or chemical.
MISSED THIS? *Read Section 1.4*

a. colorless **b.** flammable
c. liquid at room temperature **d.** density $= 0.79\ g/mL$
e. mixes with water

44. Classify each of the listed properties of ozone (a pollutant in the lower atmosphere but part of a protective shield against UV light in the upper atmosphere) as physical or chemical.

a. bluish color
b. pungent odor
c. very reactive
d. decomposes on exposure to ultraviolet light
e. gas at room temperature

45. Classify each property as physical or chemical.
MISSED THIS? *Read Section 1.4*

a. the tendency of ethyl alcohol to burn
b. the shine on silver
c. the odor of paint thinner
d. the flammability of propane gas

46. Classify each property as physical or chemical.

a. the boiling point of ethyl alcohol
b. the temperature at which dry ice evaporates
c. the tendency of iron to rust
d. the color of gold

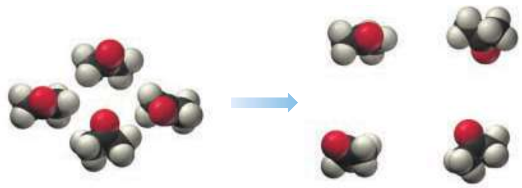
47. Classify each change as physical or chemical.
MISSED THIS? *Read Section 1.4*

a. Natural gas burns in a stove.
b. The liquid propane in a gas grill evaporates because the valve was left open.
c. The liquid propane in a gas grill burns in a flame.
d. A bicycle frame rusts on repeated exposure to air and water.

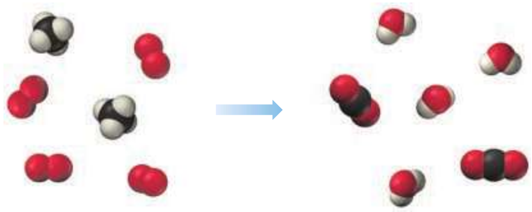
48. Classify each change as physical or chemical.

a. Sugar burns when heated in a skillet.
b. Sugar dissolves in water.
c. A platinum ring becomes dull because of continued abrasion.
d. A silver surface becomes tarnished after exposure to air for a long period of time.

49. Based on the molecular diagram, classify each change as physical or chemical. **MISSED THIS?** *Read Section 1.4*



(a)



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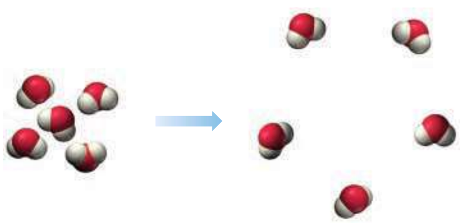
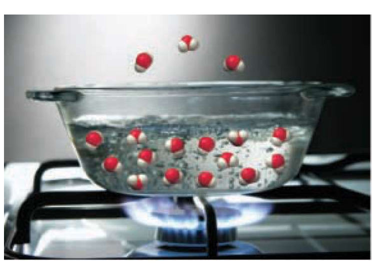


(c)

50. Based on the molecular diagram, classify each change as physical or chemical.

(a)

(b)



(c)

Units in Measurement

51. Convert each temperature.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

a. 32 °F to °C (temperature at which water freezes)
b. 77 K to °F (temperature of liquid nitrogen)
c. –109 °F to °C (temperature of dry ice)
d. 98.6 °F to K (body temperature)

52. Convert each temperature.

a. 212 °F to °C (temperature of boiling water at sea level)
b. 22 °C to K (approximate room temperature)
c. 0.00 K to °F (coldest temperature possible, also known as absolute zero)
d. 2.735 K to °C (average temperature of the universe as measured from background black body radiation)

53. The coldest ground-level temperature ever measured on Earth is –128.6 °F, recorded on July 21, 1983, in Antarctica. Convert that temperature to °C and K.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

54. The warmest temperature ever measured in the United States is 134 °F, recorded on July 10, 1913, in Death Valley, California. Convert that temperature to °C and K.

55. Use the prefix multipliers to express each measurement without exponents. **MISSED THIS?** *Read Section 1.6; Watch KCV 1.6*

a. 1.2×10^{-9} m
b. 22×10^{-15} s
c. 1.5×10^{9} g
d. 3.5×10^{6} L

56. Use prefix multipliers to express each measurement without exponents.

a. 38.8×10^{5} g
b. 55.2×10^{-10} s
c. 23.4×10^{11} m
d. 87.9×10^{-7} L

57. Use scientific notation to express each quantity with only base units (no prefix multipliers).
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

a. 4.5 ns
b. 18 fs
c. 128 pm
d. 35 μm

58. Use scientific notation to express each quantity with only base units (no prefix multipliers).

a. 35 μL
b. 225 Mm
c. 133 Tg
d. 1.5 cg

59. Complete the table.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

a.	1245 kg	1.245×10^{6} g	1.245×10^{9} mg
b.	515 km	_______dm	_______cm
c.	122.355 s	_______ms	_______ks
d.	3.345 kJ	_______J	_______mJ

60. Complete the table.

a.	355 km/s	_______cm/s	_______m/ms
b.	1228 g/L	_______g/mL	_______kg/mL
c.	554 mK/s	_______K/s	_______μK/ms
d.	2.554 mg/mL	_______g/L	_______μg/mL

61. Express the quantity 254,998 m in each unit.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*
a. km
b. Mm
c. mm
d. cm

62. Express the quantity 556.2×10^{-12} s in each unit.
a. ms **b.** ns **c.** ps **d.** fs

63. How many 1-cm squares would it take to construct a square that is 1 m on each side? **MISSED THIS?** *Read Section 1.6*

64. How many 1-cm cubes would it take to construct a cube that is 4 cm on edge?

Density

65. A new penny has a mass of 2.49 g and a volume of 0.349 cm^3. Is the penny made of pure copper? Explain your answer.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

66. A titanium bicycle frame displaces 0.314 L of water and has a mass of 1.41 kg. What is the density of the titanium in g/cm^3?

67. Glycerol is a syrupy liquid often used in cosmetics and soaps. A 3.25 L sample of pure glycerol has a mass of 4.10×10^3 g. What is the density of glycerol in g/cm^3?
MISSED THIS? *Read Section 1.6; Watch KCV 1.6*

68. A supposedly gold nugget displaces 19.3 mL of water and has a mass of 371 g. Could the nugget be made of gold?

69. Ethylene glycol (antifreeze) has a density of 1.11 g/cm^3.
MISSED THIS? *Read Section 1.6; Watch KCV 1.6, IWE 1.10*
a. What is the mass in g of 417 mL of ethylene glycol?
b. What is the volume in L of 4.1 kg of ethylene glycol?

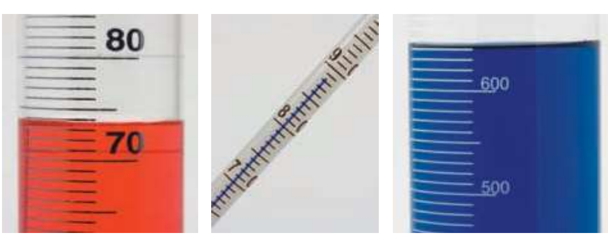
70. Acetone (nail polish remover) has a density of 0.7857 g/cm^3.
a. What is the mass in g of 28.56 mL of acetone?
b. What is the volume in mL of 6.54 g of acetone?

71. A small airplane takes on 245 L of fuel. If the density of the fuel is 0.821 g/mL, what mass of fuel has the airplane taken on?
MISSED THIS? *Read Section 1.6; Watch KCV 1.6, IWE 1.10*

72. Human fat has a density of 0.918 g/cm^3. How much volume (in cm^3) is gained by a person who gains 10.0 lb of pure fat?

The Reliability of a Measurement and Significant Figures

73. Read each measurement to the correct number of significant figures. Laboratory glassware should always be read from the bottom of the meniscus. **MISSED THIS?** *Read Section 1.7*



(a) (b) (c)

74. Read each measurement to the correct number of significant figures. Laboratory glassware should always be read from the bottom of the meniscus. Digital balances normally display mass to the correct number of significant figures for that particular balance.

(a) (b) (c)

75. For each number, underline the zeroes that are significant and draw an **x** through the zeroes that are not.
MISSED THIS? *Read Section 1.7; Watch KCV 1.6, IWE 1.5*
a. 1,050,501 km
b. 0.0020 m
c. 0.000000000000002 s
d. 0.001090 cm

76. For each number, underline the zeroes that are significant and draw an **x** through the zeroes that are not.
a. 180,701 mi
b. 0.001040 m
c. 0.005710 km
d. 90,201 m

77. How many significant figures are in each number?
MISSED THIS? *Read Section 1.7; Watch KCV 1.6, IWE 1.5*
a. 0.000312 m
b. 312,000 s
c. 3.12×10^5 km
d. 13,127 s
e. 2000

78. How many significant figures are in each number?
a. 0.1111 s
b. 0.007 m
c. 108,700 km
d. 1.563300×10^{11} m
e. 30,800

79. Which numbers are exact (and therefore have an unlimited number of significant figures)?
MISSED THIS? *Read Section 1.7; Watch KCV 1.6, IWE 1.5*
a. $\pi = 3.14$
b. 12 in = 1 ft
c. EPA gas mileage rating of 26 miles per gallon
d. 1 gross = 144

80. Indicate the number of significant figures in each number. If the number is an exact number, indicate an unlimited number of significant figures.
a. 325,365,189 (July 4, 2017 U.S. population)
b. 2.54 cm = 1 in
c. 11.4 g/cm^3 (density of lead)
d. 12 = 1 dozen

81. Round each number to four significant figures.
MISSED THIS? *Read Section 1.7; Watch KCV 1.7*
a. 156.852
b. 156.842
c. 156.849
d. 156.899

82. Round each number to three significant figures.
a. 79,845.82
b. 1.548937×10^7
c. 2.3499999995
d. 0.000045389

Significant Figures in Calculations

83. Calculate to the correct number of significant figures. **MISSED THIS?** *Read Section 1.7; Watch KCVs 1.6, 1.7, IWEs 1.5, 1.6*
a. $9.15 \div 4.970$
b. $1.54 \times 0.03060 \times 0.69$
c. $27.5 \times 1.82 \div 100.04$
d. $(2.290 \times 10^6) \div (6.7 \times 10^4)$

84. Calculate to the correct number of significant figures.
a. $89.3 \times 77.0 \times 0.08$
b. $(5.01 \times 10^5) \div (7.8 \times 10^2)$
c. $4.005 \times 74 \times 0.007$
d. $453 \div 2.031$

85. Calculate to the correct number of significant figures. **MISSED THIS?** *Read Section 1.7; Watch KCVs 1.6, 1.7, IWEs 1.5, 1.6*
a. $43.7 - 2.341$
b. $17.6 + 2.838 + 2.3 + 110.77$
c. $19.6 + 58.33 - 4.974$
d. $5.99 - 5.572$

86. Calculate to the correct number of significant figures.
a. $0.004 + 0.09879$
b. $1239.3 + 9.73 + 3.42$
c. $2.4 - 1.777$
d. $532 + 7.3 - 48.523$

87. Calculate to the correct number of significant figures. **MISSED THIS?** *Read Section 1.7; Watch KCVs 1.6, 1.7, IWEs 1.5, 1.6*
a. $(24.6681 \times 2.38) + 332.58$
b. $(85.3 - 21.489) \div 0.0059$
c. $(512 \div 986.7) + 5.44$
d. $[(28.7 \times 10^5) \div 48.533] + 144.99$

88. Calculate to the correct number of significant figures.
a. $[(1.7 \times 10^6) \div (2.63 \times 10^5)] + 7.33$
b. $(568.99 - 232.1) \div 5.3$
c. $(9443 + 45 - 9.9) \times 8.1 \times 10^6$
d. $(3.14 \times 2.4367) - 2.34$

89. A flask containing 11.7 mL of a liquid weighs 132.8 g with the liquid in the flask and 124.1 g when empty. Calculate the density of the liquid in g/mL to the correct number of significant digits. **MISSED THIS?** *Read Section 1.6; Watch KCV 1.7, IWE 1.6*

90. A flask containing 9.55 mL of a liquid weighs 157.2 g with the liquid in the flask and 148.4 g when empty. Calculate the density of the liquid in g/mL to the correct number of significant digits.

Unit Conversions

91. Perform each unit conversion.
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.8*
a. 27.8 L to cm^3
b. 1898 mg to kg
c. 198 km to cm

92. Perform each unit conversion.
a. 28.9 nm to μm
b. 1432 cm^3 to L
c. 1211 Tm to Gm

93. Perform each unit conversion.
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.8*
a. 154 cm to in
b. 3.14 kg to g
c. 3.5 L to qt
d. 109 mm to in

94. Perform each unit conversion.
a. 1.4 in to mm
b. 116 ft to cm
c. 1845 kg to lb
d. 815 yd to km

95. A runner wants to run 10.0 km. Her running pace is 7.5 mi per hour. How many minutes must she run?
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.8*

96. A cyclist rides at an average speed of 18 mi per hour. If she wants to bike 212 km, how long (in hours) must she ride?

97. A certain European automobile has a gas mileage of 17 km/L. What is the gas mileage in miles per gallon?
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.8*

98. A gas can holds 5.0 gal of gasoline. Express this quantity in cm^3.

99. A house has an area of 195 m^2. What is its area in each unit?
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.9*
a. km^2
b. dm^2
c. cm^2

100. A bedroom has a volume of 115 m^3. What is its volume in each unit?
a. km^3
b. dm^3
c. cm^3

101. The average U.S. farm occupies 435 acres. How many square miles is this? (1 acre = 43,560 ft^2, 1 mile = 5280 ft)
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.9*

102. Total U.S. farmland occupies 954 million acres. How many square miles is this? (1 acre = 43,560 ft^2, 1 mi = 5280 ft). Total U.S. land area is 3.537 million square miles. What percentage of U.S. land is farmland?

103. An acetaminophen suspension for infants contains 80 mg/0.80 mL suspension. The recommended dose is 15 mg/kg body weight. How many mL of this suspension should be given to an infant weighing 14 lb? (Assume two significant figures.)
MISSED THIS? *Read Section 1.8; Watch KCV 1.8, IWE 1.8*

104. An ibuprofen suspension for infants contains 100 mg/5.0 mL suspension. The recommended dose is 10 mg/kg body weight. How many mL of this suspension should be given to an infant weighing 18 lb? (Assume two significant figures.)

CUMULATIVE PROBLEMS

105. There are exactly 60 seconds in a minute, exactly 60 minutes in an hour, exactly 24 hours in a mean solar day, and 365.24 solar days in a solar year. How many seconds are in a solar year? Give your answer with the correct number of significant figures.

106. Determine the number of picoseconds in 2.0 hours.

107. Classify each property as intensive or extensive.
a. volume
b. boiling point
c. temperature
d. electrical conductivity
e. energy

108. At what temperatures are the readings on the Fahrenheit and Celsius thermometers the same?

109. Suppose you design a new thermometer called the X thermometer. On the X scale the boiling point of water is 130 °X, and the freezing point of water is 10 °X. At what temperature are the readings on the Fahrenheit and X thermometers the same?

110. On a new Jekyll temperature scale, water freezes at 17 °J and boils at 97 °J. On another new temperature scale, the Hyde scale, water freezes at 0 °H and boils at 120 °H. If methyl alcohol boils at 84 °H, what is its boiling point on the Jekyll scale?

111. Force is defined as mass times acceleration. Starting with SI base units, derive a unit for force. Using SI prefixes, suggest a convenient unit for the force resulting from a collision with a 10-ton trailer truck moving at 55 mi per hour and for the force resulting from the collision of a molecule of mass around 10^{-20} kg moving almost at the speed of light (3×10^8 m/s) with the wall of its container. (Assume a 1-second deceleration time for both collisions.)

112. A temperature measurement of 25 °C has three significant figures, while a temperature measurement of −196 °C has only two significant figures. Explain.

113. Do each calculation without your calculator and give the answers to the correct number of significant figures.

a. $1.76 \times 10^{-3}/8.0 \times 10^{2}$

b. $1.87 \times 10^{-2} + 2 \times 10^{-4} - 3.0 \times 10^{-3}$

c. $[(1.36 \times 10^{5})(0.000322)/0.082](129.2)$

114. The value of the euro was recently \$1.15 U.S., and the price of 1 liter of gasoline in France is 1.42 euro. What is the price of 1 gallon of gasoline in U.S. dollars in France?

115. A thief uses a can of sand to replace a solid gold cylinder that sits on a weight-sensitive, alarmed pedestal. The can of sand and the gold cylinder have exactly the same dimensions (length = 22 and radius = 3.8 cm).

a. Calculate the mass of each cylinder (ignore the mass of the can itself). (density of gold = 19.3 g/cm^3, density of sand = 3.00 g/cm^3)

b. Does the thief set off the alarm? Explain.

116. The proton has a radius of approximately 1.0×10^{-13} cm and a mass of 1.7×10^{-24} g. Determine the density of a proton. For a sphere, $V = (4/3)\pi r^3$.

117. The density of titanium is 4.51 g/cm^3. What is the volume (in cubic inches) of 3.5 lb of titanium?

118. The density of iron is 7.86 g/cm^3. What is its density in pounds per cubic inch (lb/in^3)?

119. A steel cylinder has a length of 2.16 in, a radius of 0.22 in, and a mass of 41 g. What is the density of the steel in g/cm^3?

120. A solid aluminum sphere has a mass of 85 g. Use the density of aluminum to find the radius of the sphere in inches.

121. A backyard swimming pool holds 185 cubic yards (yd^3) of water. What is the mass of the water in pounds?

122. An iceberg has a volume of 7655 ft^2. What is the mass of the ice (in kg) composing the iceberg (at 0 °C)?

123. The Toyota Prius, a hybrid electric vehicle, has an EPA gas mileage rating of 52 mi/gal in the city. How many kilometers can the Prius travel on 15 L of gasoline?

124. The Honda Insight, a hybrid electric vehicle, has an EPA gas mileage rating of 41 mi/gal in the city. How many kilometers can the Insight travel on the amount of gasoline that would fit in a soda can? The volume of a soda can is 355 mL.

125. The single proton that forms the nucleus of the hydrogen atom has a radius of approximately 1.0×10^{-13} cm. The hydrogen atom itself has a radius of approximately 52.9 pm. What fraction of the space within the atom is occupied by the nucleus?

126. A sample of gaseous neon atoms at atmospheric pressure and 0 °C contains 2.69×10^{22} atoms per liter. The atomic radius of neon is 69 pm. What fraction of the space do the atoms themselves occupy? What does this reveal about the separation between atoms in the gaseous phase?

127. The diameter of a hydrogen atom is 212 pm. Find the length in kilometers of a row of 6.02×10^{23} hydrogen atoms. The diameter of a ping pong ball is 4.0 cm. Find the length in kilometers of a row of 6.02×10^{23} ping pong balls.

128. The world record in the men's 100-m dash is 9.58 s, and in the 100-yd dash it is 9.07 s. Find the speed in mi/hr of the runners who set these records. (Assume three significant figures for 100 m and 100 yd.)

129. Table salt contains 39.33 g of sodium per 100 g of salt. The U.S. Food and Drug Administration (FDA) recommends that adults consume less than 2.40 g of sodium per day. A particular snack mix contains 1.25 g of salt per 100 g of the mix. What mass of the snack mix can an adult consume and still be within the FDA limit? (Assume three significant figures for 100 g.)

130. Lead metal can be extracted from a mineral called galena, which contains 86.6% lead by mass. A particular ore contains 68.5% galena by mass. If the lead can be extracted with 92.5% efficiency, what mass of ore is required to make a lead sphere with a 5.00-cm radius?

131. A length of #8 copper wire (radius = 1.63 mm) has a mass of 24.0 kg and a resistance of 2.061 ohm per km (Ω/km). What is the overall resistance of the wire?

132. Rolls of aluminum foil are 304 mm wide and 0.016 mm thick. What maximum length of aluminum foil can be made from 1.10 kg of aluminum?

133. Liquid nitrogen has a density of 0.808 g/mL and boils at 77 K. Researchers often purchase liquid nitrogen in insulated 175 L tanks. The liquid vaporizes quickly to gaseous nitrogen (which has a density of 1.15 g/L at room temperature and atmospheric pressure) when the liquid is removed from the tank. Suppose that all 175 L of liquid nitrogen in a tank accidentally vaporized in a lab that measured $10.00\,\text{m} \times 10.00\,\text{m} \times 2.50\,\text{m}$. What maximum fraction of the air in the room could be displaced by the gaseous nitrogen?

134. Mercury is often used in thermometers. The mercury sits in a bulb on the bottom of the thermometer and rises up a thin capillary as the temperature rises. Suppose a mercury thermometer contains 3.380 g of mercury and has a capillary that is 0.200 mm in diameter. How far does the mercury rise in the capillary when the temperature changes from 0.0 °C to 25.0 °C? The density of mercury at these temperatures is 13.596 g/cm^3 and 13.534 g/cm^3, respectively.

CHALLENGE PROBLEMS

135. A force of 2.31×10^4 N is applied to a diver's face mask that has an area of 125 cm^2. Find the pressure in atm on the face mask.

136. The SI unit of force is the newton, derived from the base units by using the definition of force, $F = ma$. The dyne is a non-SI unit of force in which mass is measured in grams and time is measured in seconds. The relationship between the two units is 1 dyne $= 10^{-5}$ N. Find the unit of length used to define the dyne.

137. Kinetic energy can be defined as $\frac{1}{2}mv^2$ or as $\frac{3}{2}PV$. Show that the derived SI units of each of these terms are those of energy. (Pressure is force/area and force is mass × acceleration.)

138. In 1999, scientists discovered a new class of black holes with masses 100 to 10,000 times the mass of our sun that occupy less space than our moon. Suppose that one of these black holes has a mass of 1×10^3 suns and a radius equal to one-half the radius of our moon. What is the density of the black hole in g/cm^3? The radius of our sun is 7.0×10^5 km, and it has an average density of 1.4×10^3 kg/m^3. The diameter of the moon is 2.16×10^3 mi.

139. Suppose that polluted air has carbon monoxide (CO) levels of 15.0 ppm. An average human inhales about 0.50 L of air per breath and takes about 20 breaths per minute. How many milligrams of carbon monoxide does the average person inhale in an 8-hour period at this level of carbon monoxide pollution? Assume that the carbon monoxide has a density of 1.2 g/L. (*Hint*: 15.0 ppm CO means 15.0 L CO per 10^6 L air.)

140. Nanotechnology, the field of building ultrasmall structures one atom at a time, has progressed in recent years. One potential application of nanotechnology is the construction of artificial cells. The simplest cells would probably mimic red blood cells, the body's oxygen transporters. Nanocontainers, perhaps constructed of carbon, could be pumped full of oxygen and injected into a person's bloodstream. If the person needed additional oxygen—due to a heart attack perhaps, or for the purpose of space travel—these containers could slowly release oxygen into the blood, allowing tissues that would otherwise die to remain alive. Suppose that the nanocontainers were cubic and had an edge length of 25 nm.

a. What is the volume of one nanocontainer? (Ignore the thickness of the nanocontainer's wall.)

b. Suppose that each nanocontainer could contain pure oxygen pressurized to a density of 85 g/L. How many grams of oxygen could each nanocontainer contain?

c. Air typically contains about 0.28 g of oxygen per liter. An average human inhales about 0.50 L of air per breath and takes about 20 breaths per minute. How many grams of oxygen does a human inhale per hour? (Assume two significant figures.)

d. What is the minimum number of nanocontainers that a person would need in his or her bloodstream to provide 1 hour's worth of oxygen?

e. What is the minimum volume occupied by the number of nanocontainers calculated in part d? Is such a volume feasible, given that total blood volume in an adult is about 5 L?

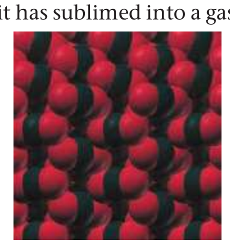
141. Approximate the percent increase in waist size that occurs when a 155-lb person gains 40.0 lb of fat. Assume that the volume of the person can be modeled by a cylinder that is 4.0 ft tall. The average density of a human is about 1.0 g/cm^3, and the density of fat is 0.918 g/cm^3.

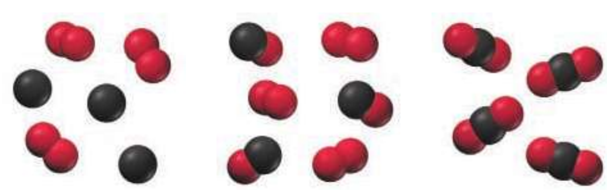
142. A box contains a mixture of small copper spheres and small lead spheres. The total volume of both metals is measured by the displacement of water to be 427 cm^3, and the total mass is 4.36 kg. What percentage of the spheres are copper?

CONCEPTUAL PROBLEMS

143. A volatile liquid (one that easily evaporates) is put into a jar, and the jar is then sealed. Does the mass of the sealed jar and its contents change upon the vaporization of the liquid?

144. The diagram shown first represents solid carbon dioxide, also known as dry ice. Which of the other diagrams best represents the dry ice after it has sublimed into a gas?





(a) (b) (c)

145. A cube has an edge length of 7 cm. If it is divided into 1-cm cubes, how many 1-cm cubes are there?

146. Substance A has a density of 1.7 g/cm^3. Substance B has a density of 1.7 kg/m^3. Without doing any calculations, determine which substance is more dense.

147. For each box, examine the blocks attached to the balances. Based on their positions and sizes, determine which block is more dense (the dark block or the lighter-colored block), or if the relative densities cannot be determined. (Think carefully about the information being shown.)

(a) (b) (c)

148. Let a triangle represent atoms of element A and a circle represent atoms of element B.

a. Draw an atomic-level view of a homogeneous mixture of elements A and B.

b. Draw an atomic view of the compound AB in a liquid state (molecules close together).

c. Draw an atomic view of the compound AB after it has undergone a physical change (such as evaporation).

d. Draw an atomic view of the compound after it has undergone a chemical change (such as decomposition of AB into A and B).

149. Identify each statement as being most like an observation, a law, or a theory.

a. All coastal areas experience two high tides and two low tides each day.

b. The tides in Earth's oceans are caused mainly by the gravitational attraction of the moon.

c. Yesterday, high tide in San Francisco Bay occurred at 2:43 A.M. and 3:07 P.M.

d. Tides are higher at the full moon and new moon than at other times of the month.

QUESTIONS FOR GROUP WORK

Active Classroom Learning

Discuss these questions with the group and record your consensus answer.

150. Using white and black circles to represent different kinds of atoms, make a drawing that accurately represents each sample of matter: a solid element, a liquid compound, and a heterogeneous mixture. Make a drawing (clearly showing *before* and *after*) depicting your liquid compound undergoing a physical change. Make a drawing depicting your solid element undergoing a chemical change.

151. Look up the measurement of the approximate thickness of a human hair.

a. Convert the measurement to an SI unit (if it isn't already).

b. Write it in scientific notation.

c. Write it without scientific notation.

d. Write it with an appropriate prefix on a base unit.

Now repeat these steps using the distance from Earth to the sun.

152. The following statements are all true.

a. Jessica's house is 5 km from the grocery store.

b. Jessica's house is 4.73 km from the grocery store.

c. Jessica's house is 4.73297 km from the grocery store.

How can all the statements be true? What does the number of digits in each statement communicate? What sort of device would Jessica need to make the measurement in each statement?

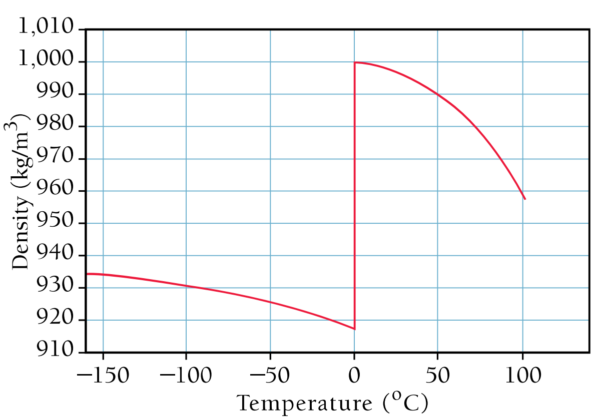
153. One inch is equal to 2.54 cm. Draw a line that is 1 in long, and mark the centimeters on the line. Draw a cube that is 1 in on each side. Draw lines on each face of the cube that are 1 cm apart. How many cubic centimeters are there in 1 in^3?

154. Convert the height of each member in your group from feet and inches to meters. Once you have your heights in meters, calculate the sum of all the heights. Use appropriate rules for significant figures at each step.

DATA INTERPRETATION AND ANALYSIS

Density of Water

155. The density of a substance can change with temperature. The graph that follows displays the density of water from −150 °C to 100 °C. Examine the graph and answer the questions.



a. Water undergoes a large change in density at 0 °C as it freezes to form ice. Calculate the percent change in density that occurs when liquid water freezes to ice at 0 °C.

(*Hint*: $\% \text{ change} = \frac{\text{final value} - \text{initial value}}{\text{initial value}} \times 100\%$)

b. Calculate the volume (in cm^3) of 54 g of water at 1 °C and the volume of the same mass of ice at −1 °C. What is the change in volume?

c. Antarctica contains 26.5 million cubic kilometers of ice. Assume the temperature of the ice is −20 °C. If all of this ice were heated to 1 °C and melted to form water, what volume of liquid water would form?

d. A 1.00-L sample of water is heated from 1 °C to 100 °C. What is the volume of the water after it is heated?

Cc ANSWERS TO CONCEPTUAL CONNECTIONS

Laws and Theories

1.1 **(b)** A law only summarizes a series of related observations; a theory gives the underlying reasons for them.

Pure Substances and Mixtures

1.2 **(a)** This image is a pure substance. More specifically, because it contains two different type of atoms bonded together, it is a pure compound.

Chemical and Physical Changes

1.3 View **(a)** best represents the water after vaporization. Vaporization is a physical change, so the molecules must remain the same before and after the change.

Energy

1.4 **(c)** Chemical energy is a type of potential energy that results from the electrostatic forces between the charged particles that compose atoms and molecules.

Temperature Scale

1.5 **(a)** The Kelvin scale has no negative temperatures because 0 Kelvin is the coldest possible temperature. Lower temperatures do not exist. Both the Celsius scale and the Fahrenheit scale have negative temperatures.

Prefix Multipliers

1.6 **(c)** The prefix micro (10^{-6}) is appropriate. The measurement would be reported as 55.7 μm.

Density

1.7 **(c)** The copper sample expands. However, because its mass remains constant while its volume increases, its density decreases.