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Introduction: Matter and Measurement

In the title of this book we refer to chemistry as the *central science*. This title reflects the fact that much of what goes on in the world around us involves chemistry. The changes that produce the brilliant colors of tree leaves in the fall, the electrical energy that powers a cell phone, the spoilage of foods left standing at room temperature, and the many ways in which our bodies use the foods we consume are all everyday examples of chemical processes.

Chemistry is the study of matter and the changes that matter undergoes. As you progress in your study, you will come to see how chemical principles operate in all aspects of our lives, from everyday activities like food preparation to more complex processes such as those that operate in the environment. We use chemical principles to understand a host of phenomena, from the role of salt in our diet to the workings of a lithium ion battery.

This first chapter provides an overview of what chemistry is about and what chemists do. The "What's Ahead" list gives an overview of the chapter organization and of some of the ideas we will consider.

1.1 | The Study of Chemistry

Chemistry is at the heart of many changes we see in the world around us, and it accounts for the myriad of different properties we see in matter. To understand how these changes and properties arise, we need to look far beneath the surfaces of our everyday observations.

▶ **THE BEAUTIFUL COLORS** that develop in trees in the fall appear when the tree ceases to produce chlorophyll, which imparts the green color to the leaves during the summer. Some of the color we see has been in the leaf all summer, and some develops from the action of sunlight on the leaf as the chlorophyll disappears.

WHAT'S AHEAD

1.1 THE STUDY OF CHEMISTRY We begin with a brief description of what chemistry is, what chemists do, and why it is useful to learn chemistry.

1.2 CLASSIFICATIONS OF MATTER Next, we examine some fundamental ways to classify matter, distinguishing between *pure substances* and *mixtures* and between *elements* and *compounds*.

1.3 PROPERTIES OF MATTER We then consider different characteristics, or *properties*, used to characterize, identify, and separate substances, distinguishing between chemical and physical properties.

1.4 UNITS OF MEASUREMENT We observe that many properties rely on quantitative measurements involving numbers and units. The units of measurement used throughout science are those of the *metric system*.

1.5 UNCERTAINTY IN MEASUREMENT We observe that the uncertainty inherent in all measured quantities is expressed by the number of *significant figures* used to report the quantity. Significant figures are also used to express the uncertainty associated with calculations involving measured quantities.

1.6 DIMENSIONAL ANALYSIS We recognize that units as well as numbers are carried through calculations and that obtaining correct units for the result of a calculation is an important way to check whether the calculation is correct.

The Atomic and Molecular Perspective of Chemistry

Chemistry is the study of the properties and behavior of matter. **Matter** is the physical material of the universe; it is anything that has mass and occupies space. A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types. This book, your body, the air you are breathing, and the clothes you are wearing are all samples of matter. We observe a tremendous variety of matter in our world, but countless experiments have shown that all matter is comprised of combinations of only about 100 substances called ***elements***. One of our major goals will be to relate the properties of matter to its composition, that is, to the particular elements it contains.

Chemistry also provides a background for understanding the properties of matter in terms of **atoms**, the almost infinitesimally small building blocks of matter. Each element is composed of a unique kind of atom. We will see that the properties of matter relate to both the kinds of atoms the matter contains (*composition*) and the arrangements of these atoms (*structure*).

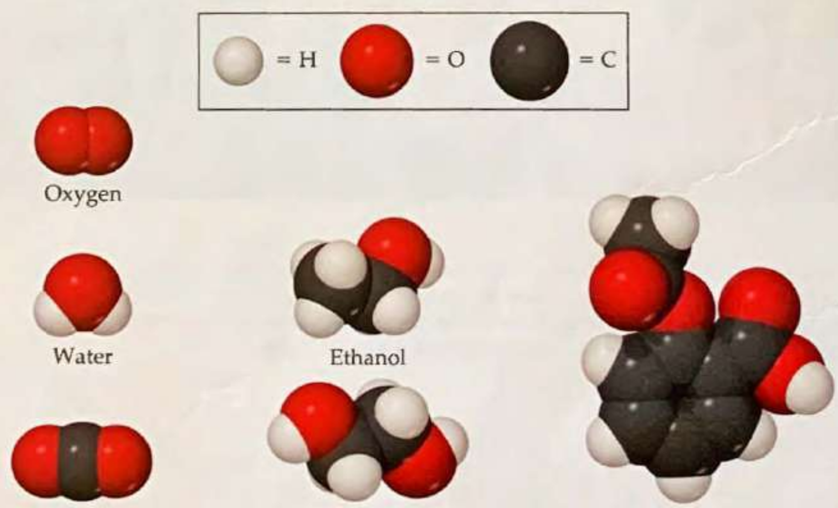
In **molecules**, two or more atoms are joined in specific shapes. Throughout this text you will see molecules represented using colored spheres to show how the atoms are connected (▼ **Figure 1.1**). The color provides a convenient way to distinguish between atoms of different elements. For example, notice that the molecules of ethanol and ethylene glycol in Figure 1.1 have different compositions and structures. Ethanol contains one oxygen atom, depicted by one red sphere. In contrast, ethylene glycol contains two oxygen atoms.

Even apparently minor differences in the composition or structure of molecules can cause profound differences in properties. For example, let's compare ethanol and ethylene glycol, which appear in Figure 1.1 to be quite similar. Ethanol is the alcohol in beverages such as beer and wine, whereas ethylene glycol is a viscous liquid used as automobile antifreeze. The properties of these two substances differ in many ways, as do their biological activities. Ethanol is consumed throughout the world, but you should *never* consume ethylene glycol because it is highly toxic. One of the challenges chemists undertake is to alter the composition or structure of molecules in a controlled way, creating new substances with different properties. For example, the common drug aspirin, shown in Figure 1.1, was first synthesized in 1897 in a successful attempt to improve on a natural product extracted from willow bark that had long been used to alleviate pain.

Every change in the observable world—from boiling water to the changes that occur as our bodies combat invading viruses—has its basis in the world of atoms and molecules.

GO FIGURE

Which of the molecules in the figure has the most carbon atoms? How many are there in that molecule?



▲ **Figure 1.1 Molecular models.** The white, black, and red spheres represent atoms of hydrogen, carbon, and oxygen, respectively.

Thus, as we proceed with our study of chemistry, we will find ourselves thinking in two realms: the *macroscopic* realm of ordinary-sized objects (*macro* = large) and the *submicroscopic* realm of atoms and molecules. We make our observations in the macroscopic world, but to understand that world, we must visualize how atoms and molecules behave at the submicroscopic level. Chemistry is the science that seeks to understand the properties and behavior of matter by studying the properties and behavior of atoms and molecules.

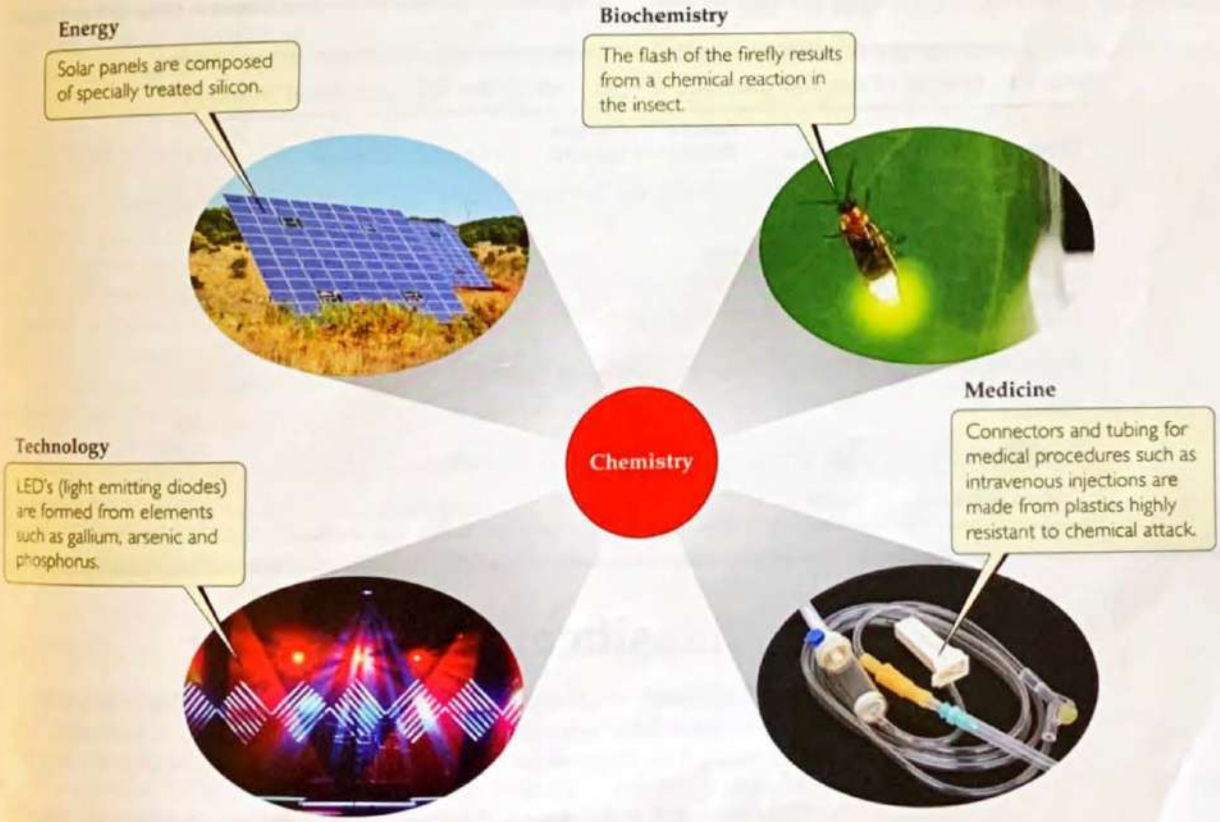
Give It Some Thought

(a) Approximately how many elements are there?
(b) What submicroscopic particles are the building blocks of matter?

Why Study Chemistry?

Chemistry lies near the heart of many matters of public concern, such as improvement of health care, conservation of natural resources, protection of the environment, and the supply of energy needed to keep society running. Using chemistry, we have discovered and continually improved upon pharmaceuticals, fertilizers and pesticides, plastics, solar panels, LEDs, and building materials. We have also discovered that some chemicals are potentially harmful to our health or the environment. This means that we must be sure that the materials with which we come into contact are safe. As a citizen and consumer, it is in your best interest to understand the effects, both positive and negative, that chemicals can have, and to arrive at a balanced outlook regarding their uses.

You may be studying chemistry because it is an essential part of your curriculum. Your major might be chemistry, or it could be biology, engineering, pharmacy, agriculture, geology, or some other field. Chemistry is central to a fundamental understanding of governing principles in many science-related fields. For example, our interactions with the material world raise basic questions about the materials around us. ▼ **Figure 1.2** illustrates how chemistry is central to several different realms of modern life.



▲ **Figure 1.2 Chemistry is central to our understanding of the world around us.**

Chemistry Put to Work

Chemistry and the Chemical Industry

Chemistry is all around us. Many people are familiar with household chemicals, particularly kitchen chemicals such as those shown in ▶ **Figure 1.3**. However, few realize the size and importance of the chemical industry. Worldwide sales of chemicals and related products manufactured in the United States total approximately $585 billion annually. Sales of pharmaceuticals total another $180 billion. The chemical industry employs more than 10% of all scientists and engineers and is a major contributor to the U.S. economy.

Vast amounts of industrial chemicals are produced each year. ▼ **Table 1.1** lists several of the chemicals produced in highest volumes in the United States. Notice that they all serve as raw materials for a variety of uses, including the manufacture and processing of metals, plastics, fertilizers, and other goods.

Who are chemists, and what do they do? People who have degrees in chemistry hold a variety of positions in industry, government, and academia. Those in industry work as laboratory chemists, developing new products (research and development); analyzing materials (quality control); or assisting customers in using products (sales and service). Those with more experience or training may work as managers or company directors. Chemists are important members of the scientific workforce in government (the National Institutes of Health, Department of Energy, and Environmental Protection Agency all employ chemists) and at universities. A chemistry degree is also good preparation for careers in teaching, medicine, biomedical research, information science, environmental work, technical sales, government regulatory agencies, and patent law.

Fundamentally, chemists do three things: (1) make new types of matter: materials, substances, or combinations of substances with desired properties; (2) measure the properties of matter; and (3) develop models that explain and/or predict the properties of matter. One chemist, for example, may work in the laboratory to discover new drugs. Another may concentrate on the development of new instrumentation to measure properties of matter at the atomic level. Other chemists may use existing materials and methods to understand how pollutants are transported in the environment or how drugs are processed in the body. Yet another chemist will develop theory, write computer code, and run computer simulations to understand how molecules move and react. The collective chemical enterprise is a rich mix of all of these activities.

▲ Figure 1.3 **Common chemicals employed in home food production.**

Table 1.1 Several of the Top Chemicals Produced by the U.S. Chemical Industry*

Chemical	Formula	Annual Production (Billions of Pounds)	Principal End Uses
Sulfuric acid	H_2SO_4	70	Fertilizers, chemical manufacturing
Ethylene	C_2H_4	50	Plastics, antifreeze
Lime	CaO	45	Paper, cement, steel
Propylene	C_3H_6	35	Plastics
Ammonia	NH_3	18	Fertilizers
Chlorine	Cl_2	21	Bleaches, plastics, water purification
Phosphoric acid	H_3PO_4	20	Fertilizers
Sodium hydroxide	$NaOH$	16	Aluminum production, soap

1.2 | Classifications of Matter

Let's begin our study of chemistry by examining two fundamental ways in which matter is classified. Matter is typically characterized by (1) its physical state (gas, liquid, or solid) and (2) its composition (whether it is an element, a *compound*, or a *mixture*).

*Data from Chemical & Engineering News, July 2, 2007, pp. 57, 60, American Chemical Society; data online from U.S. Geological Survey.

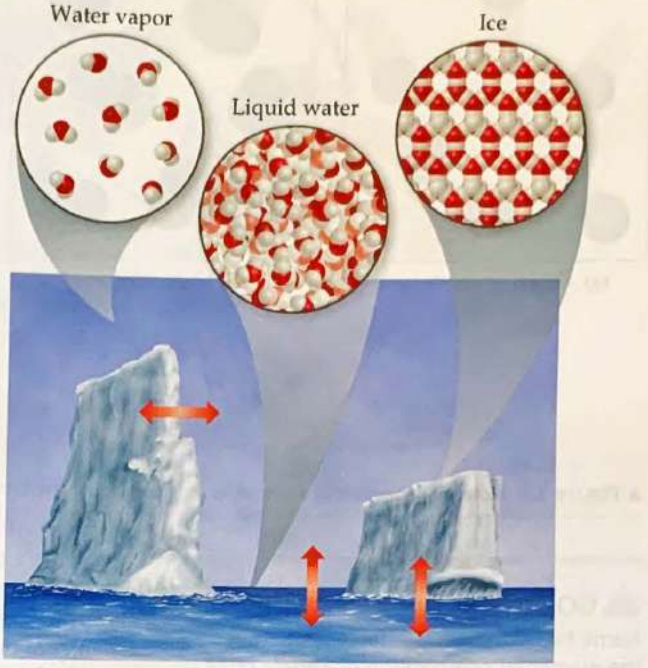
States of Matter

A sample of matter can be a gas, a liquid, or a solid. These three forms, called the **states of matter**, differ in some of their observable properties. A **gas** (also known as *vapor*) has no fixed volume or shape; rather, it uniformly fills its container. A gas can be compressed to occupy a smaller volume, or it can expand to occupy a larger one. A **liquid** has a distinct volume independent of its container, and assumes the shape of the portion of the container it occupies. A **solid** has both a definite shape and a definite volume. Neither liquids nor solids can be compressed to any appreciable extent.

The properties of the states of matter can be understood on the molecular level (▶ **Figure 1.4**). In a gas the molecules are far apart and moving at high speeds, colliding repeatedly with one another and with the walls of the container. Compressing a gas decreases the amount of space between molecules and increases the frequency of collisions between molecules but does not alter the size or shape of the molecules. In a liquid, the molecules are packed closely together but still move rapidly. The rapid movement allows the molecules to slide over one another; thus, a liquid pours easily. In a solid the molecules are held tightly together, usually in definite arrangements in which the molecules can wiggle only slightly in their otherwise fixed positions. Thus, the distances between molecules are similar in the liquid and solid states, but the two states differ in how free the molecules are to move around. Changes in temperature and/or pressure can lead to conversion from one state of matter to another, illustrated by such familiar processes as ice melting or water vapor condensing.

GO FIGURE

In which form of water are the water molecules farthest apart?



▲ **Figure 1.4 The three physical states of water—water vapor, liquid water, and ice.** We see the liquid and solid states but cannot see the gas (vapor) state. The red arrows show that the three states of matter interconvert.

Pure Substances

Most forms of matter we encounter—the air we breathe (a gas), the gasoline we burn in our cars (a liquid), and the sidewalk we walk on (a solid)—are not chemically pure. We can, however, separate these forms of matter into pure substances. A **pure substance** (usually referred to simply as a *substance*) is matter that has distinct properties and a composition that does not vary from sample to sample. Water and table salt (sodium chloride) are examples of pure substances.

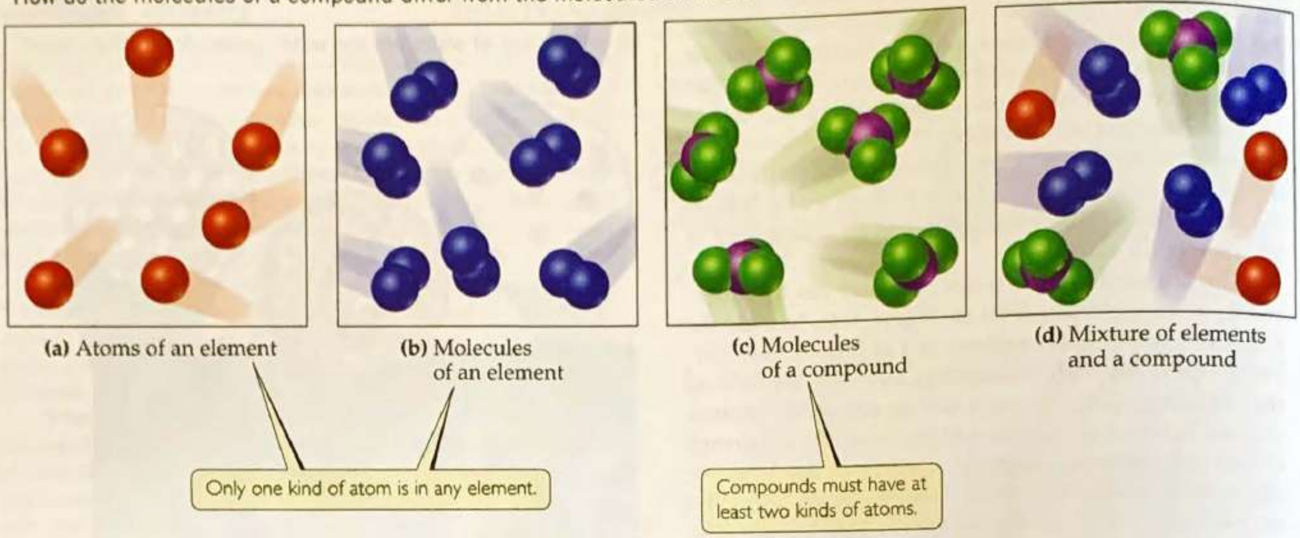
All substances are either elements or compounds. **Elements** are substances that cannot be decomposed into simpler substances. On the molecular level, each element is composed of only one kind of atom [**Figure 1.5(a and b)**]. **Compounds** are substances composed of two or more elements; they contain two or more kinds of atoms [Figure 1.5(c)]. Water, for example, is a compound composed of two elements: hydrogen and oxygen. Figure 1.5(d) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each substance retains its chemical identity.

Elements

Currently, 118 elements are known, though they vary widely in abundance. Hydrogen constitutes about 74% of the mass in the Milky Way galaxy, and helium constitutes 24%. Closer to home, only five elements—oxygen, silicon, aluminum, iron, and calcium—account for over 90% of Earth's crust (including oceans and atmosphere), and only three—oxygen, carbon, and hydrogen—account for over 90% of the mass of the human body (**Figure 1.6**).

GO FIGURE

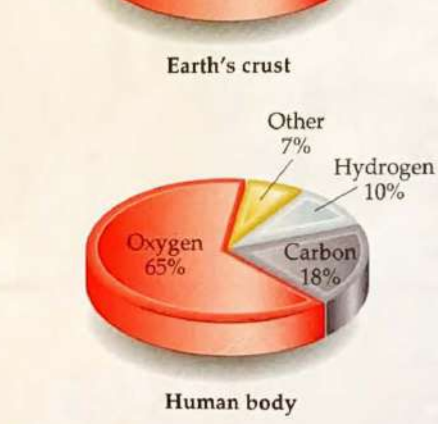
How do the molecules of a compound differ from the molecules of an element?



▲ Figure 1.5 **Molecular comparison of elements, compounds, and mixtures.**

GO FIGURE

Name two significant differences between the elemental composition of Earth's crust and the elemental composition of the human body.



▲ Figure 1.6 **Relative abundances of elements.*** Elements in percent by mass in Earth's crust (including oceans and atmosphere) and the human body.

▼ **Table 1.2** lists some common elements, along with the chemical *symbols* used to denote them. The symbol for each element consists of one or two letters, with the first letter capitalized. These symbols are derived mostly from the English names of the elements, but sometimes they are derived from a foreign name instead (last column in Table 1.2). You will need to know these symbols and learn others as we encounter them in the text.

All of the known elements and their symbols are listed on the front inside cover of this text in a table known as the *periodic table*. In the periodic table the elements are arranged in columns so that closely related elements are grouped together. We describe the periodic table in more detail in Section 2.5 and consider the periodically repeating properties of the elements in Chapter 7.

Compounds

Most elements can interact with other elements to form compounds. For example, when hydrogen gas burns in oxygen gas, the elements hydrogen and oxygen combine to form the compound water. Conversely, water can be decomposed into its elements by passing an electrical current through it (▶ **Figure 1.7**).

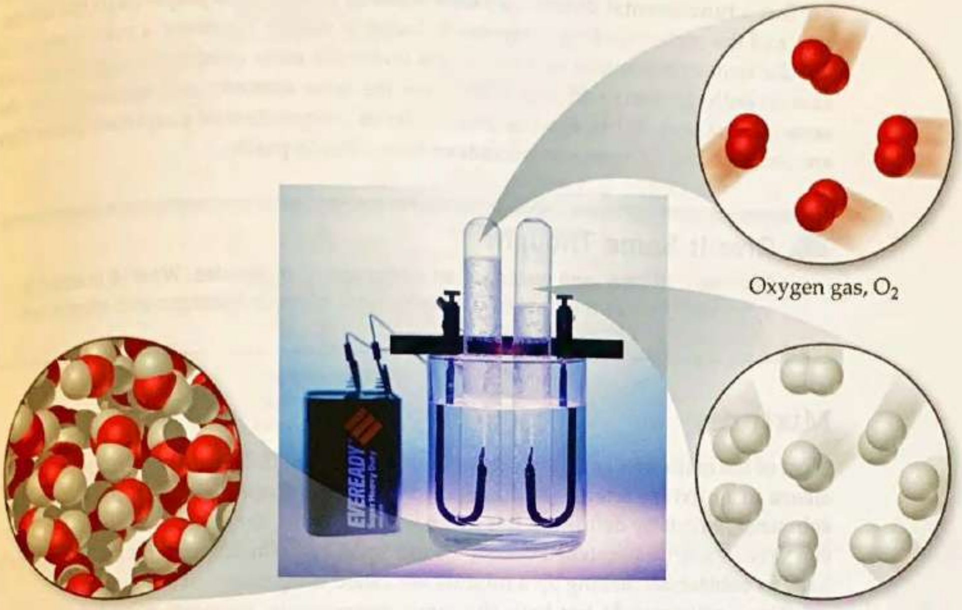
Table 1.2 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from *cuprum*)
Fluorine	F	Bromine	Br	Iron	Fe (from *ferrum*)
Hydrogen	H	Calcium	Ca	Lead	Pb (from *plumbum*)
Iodine	I	Chlorine	Cl	Mercury	Hg (from *hydrargyrum*)
Nitrogen	N	Helium	He	Potassium	K (from *kalium*)
Oxygen	O	Lithium	Li	Silver	Ag (from *argentum*)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from *natrium*)
Sulfur	S	Silicon	Si	Tin	Sn (from *stannum*)

*U.S. Geological Survey Circular 285, U.S Department of the Interior.

GO FIGURE

How are the relative gas volumes collected in the two tubes related to the relative number of gas molecules in the tubes?



▲ **Figure 1.7 Electrolysis of water.** Water decomposes into its component elements, hydrogen and oxygen, when an electrical current is passed through it. The volume of hydrogen, collected in the right test tube, is twice the volume of oxygen.

Pure water, regardless of its source, consists of 11% hydrogen and 89% oxygen by mass. This macroscopic composition corresponds to the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom:

 Hydrogen atom (written H)  Oxygen atom (written O)  Water molecule (written H_2O)

The elements hydrogen and oxygen themselves exist naturally as diatomic (two-atom) molecules:

Oxygen molecule  (written O_2)

Hydrogen molecule  (written H_2)

As seen in ▼ **Table 1.3**, the properties of water bear no resemblance to the properties of its component elements. Hydrogen, oxygen, and water are each a unique substance, a consequence of the uniqueness of their respective molecules.

Table 1.3 **Comparison of Water, Hydrogen, and Oxygen**

	Water	Hydrogen	Oxygen
State[a]	Liquid	Gas	Gas
Normal boiling point	100 °C	−253 °C	−183 °C
Density[a]	1000 g/L	0.084 g/L	1.33 g/L
Flammable	No	Yes	No

[a]At room temperature and atmospheric pressure.

The observation that the elemental composition of a compound is always the same is known as the **law of constant composition** (or the **law of definite proportions**). French chemist Joseph Louis Proust (1754–1826) first stated the law in about 1800. Although this law has been known for 200 years, the belief persists among some people that a fundamental difference exists between compounds prepared in the laboratory and the corresponding compounds found in nature. However, a pure compound has the same composition and properties under the same conditions regardless of its source. Both chemists and nature must use the same elements and operate under the same natural laws. When two materials differ in composition or properties, either they are composed of different compounds or they differ in purity.

Give It Some Thought

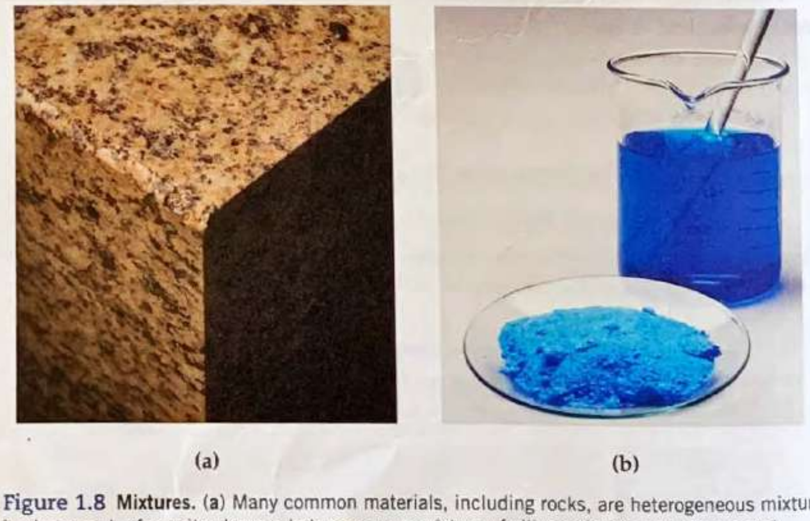
Hydrogen, oxygen, and water are all composed of molecules. What is it about a molecule of water that makes it a compound, whereas hydrogen and oxygen are elements?

Mixtures

Most of the matter we encounter consists of mixtures of different substances. Each substance in a mixture retains its chemical identity and properties. In contrast to a pure substance, which by definition has a fixed composition, the composition of a mixture can vary. A cup of sweetened coffee, for example, can contain either a little sugar or a lot. The substances making up a mixture are called *components* of the mixture.

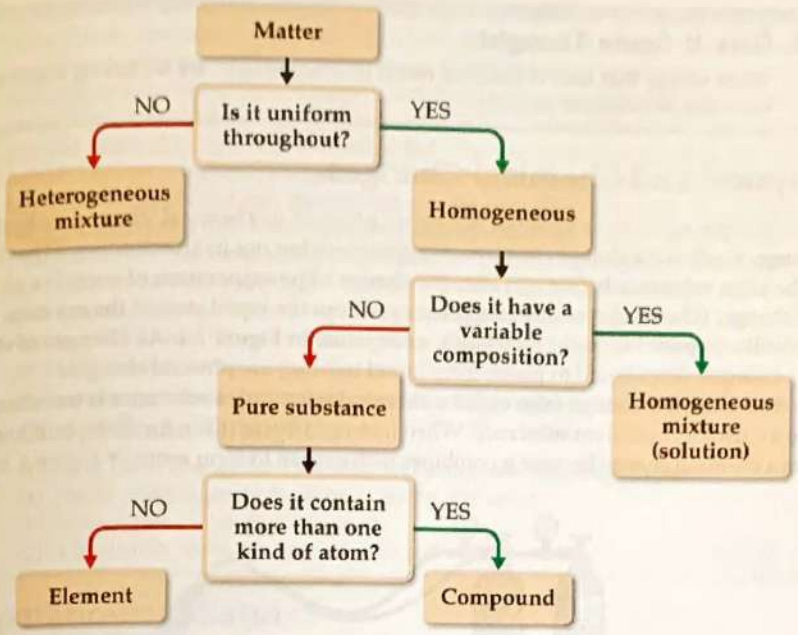
Some mixtures do not have the same composition, properties, and appearance throughout. Rocks and wood, for example, vary in texture and appearance in any typical sample. Such mixtures are *heterogeneous* [▼ **Figure 1.8(a)**]. Mixtures that are uniform throughout are *homogeneous*. Air is a homogeneous mixture of nitrogen, oxygen, and smaller amounts of other gases. The nitrogen in air has all the properties of pure nitrogen because both the pure substance and the mixture contain the same nitrogen molecules. Salt, sugar, and many other substances dissolve in water to form homogeneous mixtures [Figure 1.8(**b**)]. Homogeneous mixtures are also called **solutions**. Although the term *solution* conjures an image of a liquid, solutions can be solids, liquids, or gases.

▶ **Figure 1.9** summarizes the classification of matter into elements, compounds, and mixtures.



(a) (b)

▲ **Figure 1.8 Mixtures.** (**a**) Many common materials, including rocks, are heterogeneous mixtures. This photograph of granite shows a heterogeneous mixture of silicon dioxide and other metal oxides. (**b**) Homogeneous mixtures are called solutions. Many substances, including the blue solid shown here [copper(II) sulfate], dissolve in water to form solutions.



▲ **Figure 1.9 Classification of matter.** All pure matter is classified ultimately as either an element or a compound.

SAMPLE EXERCISE 1.1 Distinguishing among Elements, Compounds, and Mixtures

"White gold" contains gold and a "white" metal, such as palladium. Two samples of white gold differ in the relative amounts of gold and palladium they contain. Both samples are uniform in composition throughout. Use Figure 1.9 to classify white gold.

SOLUTION

Because the material is uniform throughout, it is homogeneous. Because its composition differs for the two samples, it cannot be a compound. Instead, it must be a homogeneous mixture.

Practice Exercise 1

Which of the following is the correct description of a cube of material cut from the inside of an apple?

(a) It is a pure compound.
(b) It consists of a homogenous mixture of compounds.
(c) It consists of a heterogeneous mixture of compounds.
(d) It consists of a heterogeneous mixture of elements and compounds.
(e) It consists of a single compound in different states.

Practice Exercise 2

Aspirin is composed of 60.0% carbon, 4.5% hydrogen, and 35.5% oxygen by mass, regardless of its source. Use Figure 1.9 to classify aspirin.

1.3 | Properties of Matter

Every substance has unique properties. For example, the properties listed in Table 1.3 allow us to distinguish hydrogen, oxygen, and water from one another. The properties of matter can be categorized as physical or chemical. **Physical properties** can be observed without changing the identity and composition of the substance. These properties include color, odor, density, melting point, boiling point, and hardness. **Chemical properties** describe the way a substance may change, or *react*, to form other substances. A common chemical property is flammability, the ability of a substance to burn in the presence of oxygen.

Some properties, such as temperature and melting point, are *intensive properties*. **Intensive properties** do not depend on the amount of sample being examined and are particularly useful in chemistry because many intensive properties can be used to *identify* substances. **Extensive properties** depend on the amount of sample, with two examples being mass and volume. Extensive properties relate to the *amount* of substance present.

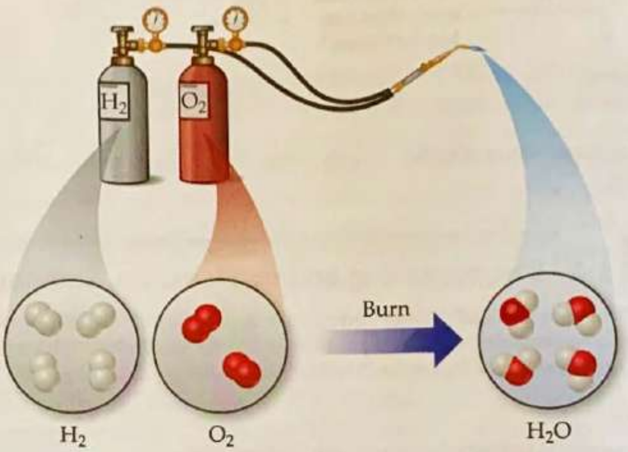
Give It Some Thought

When we say that lead is a denser metal than aluminum, are we talking about an extensive or intensive property?

Physical and Chemical Changes

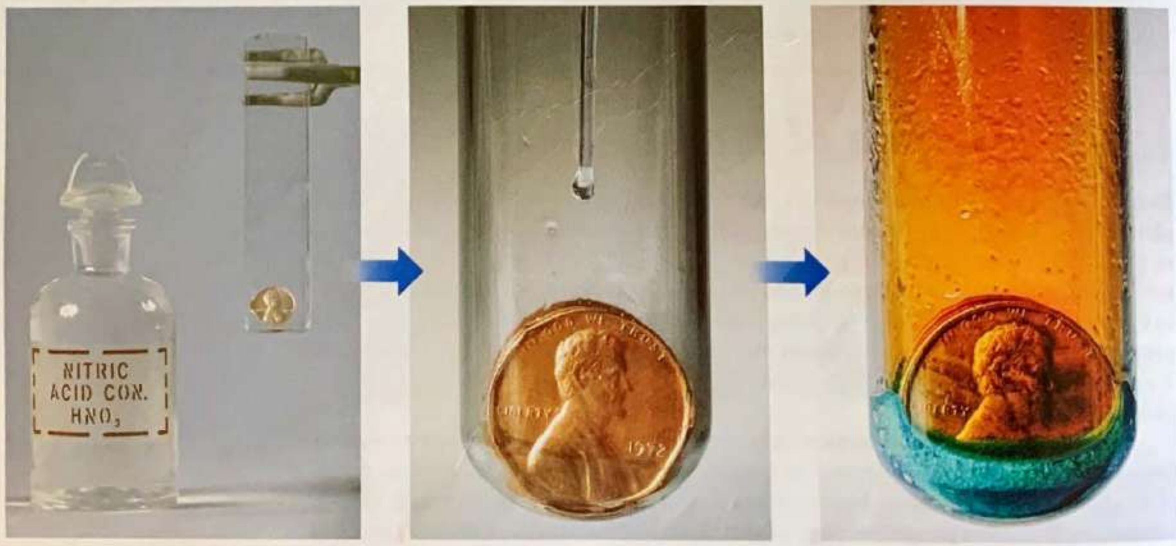
The changes substances undergo are either physical or chemical. During a **physical change**, a substance changes its physical appearance but not its composition. (That is, it is the same substance before and after the change.) The evaporation of water is a physical change. When water evaporates, it changes from the liquid state to the gas state, but it is still composed of water molecules, as depicted in Figure 1.4. All **changes of state** (for example, from liquid to gas or from liquid to solid) are physical changes.

In a **chemical change** (also called a **chemical reaction**), a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water (▼ Figure 1.10).



▲ Figure 1.10 **A chemical reaction.**

Chemical changes can be dramatic. In the account that follows, Ira Remsen, author of a popular chemistry text published in 1901, describes his first experiences with chemical reactions. The chemical reaction that he observed is shown in ▼ Figure 1.11.



▲ Figure 1.11 **The chemical reaction between a copper penny and nitric acid.** The dissolved copper produces the blue-green solution; the reddish brown gas produced is nitrogen dioxide.

While reading a textbook of chemistry, I came upon the statement "nitric acid acts upon copper," and I determined to see what this meant. Having located some nitric acid, I had only to learn what the words "act upon" meant. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table, opened a bottle labeled "nitric acid," poured some of the liquid on the copper, and prepared to make an observation. But what was this wonderful thing which I beheld? The cent was already changed, and it was no small change either. A greenish-blue liquid foamed and fumed over the cent and over the table. The air became colored dark red. How could I stop this? I tried by picking the cent up and throwing it out the window. I learned another fact: nitric acid acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and discovered nitric acid acts upon trousers. That was the most impressive experiment I have ever performed. I tell of it even now with interest. It was a revelation to me. Plainly the only way to learn about such remarkable kinds of action is to see the results, to experiment, to work in the laboratory.*

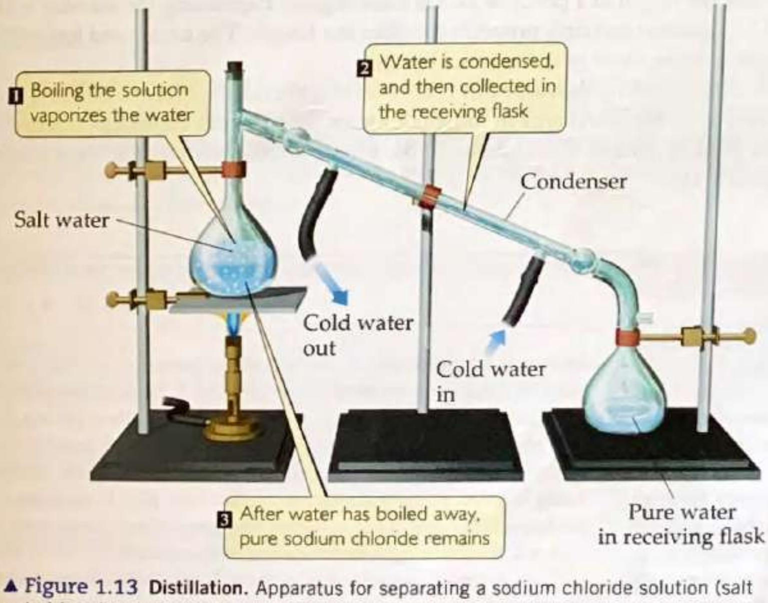
Give It Some Thought

Which of these changes are physical and which are chemical? Explain.
(a) Plants make sugar from carbon dioxide and water.
(b) Water vapor in the air forms frost.
(c) A goldsmith melts a nugget of gold and pulls it into a wire.

Separation of Mixtures

We can separate a mixture into its components by taking advantage of differences in their properties. For example, a heterogeneous mixture of iron filings and gold filings could be sorted by color into iron and gold. A less tedious approach would be to use a magnet to attract the iron filings, leaving the gold ones behind. We can also take advantage of an important chemical difference between these two metals: Many acids dissolve iron but not gold. Thus, if we put our mixture into an appropriate acid, the acid would dissolve the iron and the solid gold would be left behind. The two could then be separated by *filtration* (▶ Figure 1.12). We would have to use other chemical reactions, which we will learn about later, to transform the dissolved iron back into metal.

An important method of separating the components of a homogeneous mixture is **distillation**, a process that depends on the different abilities of substances to form gases. For example, if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous water can be converted back to a liquid on the walls of a condenser, as shown in ▼ Figure 1.13.



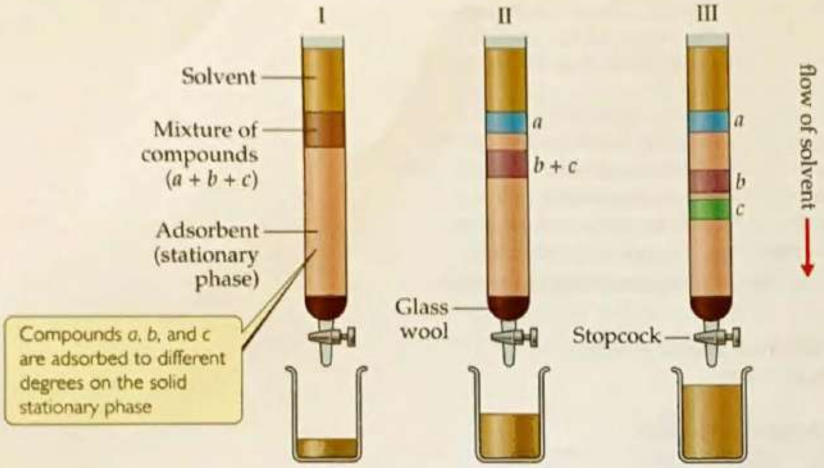
▲ Figure 1.13 **Distillation.** Apparatus for separating a sodium chloride solution (salt water) into its components.

▲ Figure 1.12 **Separation by filtration.** A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.

*Remsen, Ira, *The Principles of Theoretical Chemistry*, 1887.

GO FIGURE

Is the separation of **a**, **b**, and **c** in Figure 1.14 a physical or chemical process?



▲ Figure 1.14 **Separation of three substances using column chromatography.**

▲ Figure 1.15 **Metric units.** Metric measurements are increasingly common in the United States, as exemplified by the volume printed on this soda can in both English units (fluid ounces, fl oz) and metric units (milliliters, mL).

The differing abilities of substances to adhere to the surfaces of solids can also be used to separate mixtures. This ability is the basis of *chromatography*, a technique shown in ▲ Figure 1.14.

1.4 | Units of Measurement

Many properties of matter are *quantitative*, that is, associated with numbers. When a number represents a measured quantity, the units of that quantity must be specified. To say that the length of a pencil is 17.5 is meaningless. Expressing the number with its units, 17.5 centimeters (cm), properly specifies the length. The units used for scientific measurements are those of the **metric system.**

The metric system, developed in France during the late eighteenth century, is used as the system of measurement in most countries. The United States has traditionally used the English system, although use of the metric system has become more common (◀ Figure 1.15).

A Closer Look

The Scientific Method

Where does scientific knowledge come from? How is it acquired? How do we know it is reliable? How do scientists add to it, or modify it?

There is nothing mysterious about how scientists work. The first idea to keep in mind is that scientific knowledge is gained through observations of the natural world. A principal aim of the scientist is to organize these observations, by identifying patterns and regularity, making measurements, and associating one set of observations with another. The next step is to ask *why* nature behaves in the manner we observe. To answer this question, the scientist constructs a model, known as a **hypothesis,** to explain the observations. Initially the hypothesis is likely to be pretty tentative. There could be more than one reasonable hypothesis. If a hypothesis is correct, then certain results and observations should follow from it. In this way hypotheses can stimulate the design of experiments to learn more about the system being studied. Scientific creativity comes into play in thinking of hypotheses that are fruitful in suggesting good experiments to do, ones that will shed new light on the nature of the system.

As more information is gathered, the initial hypotheses get winnowed down. Eventually just one may stand out as most consistent with a body of accumulated evidence. We then begin to call this hypothesis a **theory**, a model that has predictive powers, and that accounts for all the available observations. A theory also generally is consistent with other, perhaps larger and more general theories. For example, a theory of what goes on inside a volcano has to be consistent with more general theories regarding heat transfer, chemistry at high temperature, and so forth.

We will be encountering many theories as we proceed through this book. Some of them have been found over and over again to be consistent with observations. However, no theory can be proven to be absolutely true. We can treat it as though it is, but there always remains a possibility that there is some respect in which a theory is wrong. A famous example is Einstein's theory of relativity. Isaac Newton's theory of mechanics yielded such precise results for the mechanical behavior of matter that no exceptions to it were found before the twentieth century. But Albert Einstein showed that Newton's theory of the nature of space and time is incorrect. Einstein's theory of relativity represented a fundamental shift in how we think of space and time. He predicted where the exceptions to predictions based on Newton's theory might be found. Although only small departures from Newton's theory were predicted, they *were* observed. Einstein's theory of relativity became accepted as the correct model. However, for most uses, Newton's laws of motion are quite accurate enough.

The overall process we have just considered, illustrated in ▶ **Figure 1.16**, is often referred to as *the scientific method*. But there is no single scientific method. Many factors play a role in advancing scientific knowledge. The one unvarying requirement is that our explanations be consistent with observations, and that they depend solely on natural phenomena.

When nature behaves in a certain way over and over again, under all sorts of different conditions, we can summarize that behavior in a **scientific law**. For example, it has been repeatedly observed that in a chemical reaction there is no change in the total mass of the materials reacting as compared with the materials that are formed; we call this observation the *Law of Conservation of Mass*. It is important to make a distinction between a theory and a scientific law. The latter simply is a statement of what always happens, to the best of our knowledge. A theory, on the other hand, is an *explanation* for what happens. If we discover some law fails to hold true, then we must assume the theory underlying that law is wrong in some way.

Related Exercises: 1.60, 1.82

▲ **Figure 1.16** The scientific method.

SI Units

In 1960 an international agreement was reached specifying a particular choice of metric units for use in scientific measurements. These preferred units are called **SI units**, after the French *Système International d'Unités*. This system has seven *base units* from which all other units are derived (▼ **Table 1.4**). In this chapter we will consider the base units for length, mass, and temperature.

Table 1.4 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s or sec
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

Give It Some Thought

The package of a fluorescent bulb for a table lamp lists the light output in terms of lumens, lm. Which of the seven SI units would you expect to be part of the definition of a lumen?

With SI units, prefixes are used to indicate decimal fractions or multiples of various units. For example, the prefix *milli-* represents a 10^{-3} fraction, one-thousandth, of a unit: A milligram (mg) is 10^{-3} gram (g), a millimeter (mm) is 10^{-3} meter (m), and so forth. ▼ **Table 1.5** presents the prefixes commonly encountered in chemistry. In using SI units and in working problems throughout this text, you must be comfortable using exponential notation. If you are unfamiliar with exponential notation or want to review it, refer to Appendix A.1.

Although non–SI units are being phased out, some are still commonly used by scientists. Whenever we first encounter a non–SI unit in the text, the SI unit will also be given. The relations between the non–SI and SI units we will use most frequently in this text appear on the back inside cover. We will discuss how to convert from one to the other in Section 1.6.

Table 1.5 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1 \text{ J} = 1 \text{ kg} \cdot \text{m}^2/\text{s}^2$ and $1 \text{ W} = 1 \text{ J/s}$.
[b]Greek letter mu, pronounced "mew."

Give It Some Thought

How many μg are there in 1 mg?

Length and Mass

The SI base unit of *length* is the meter, a distance slightly longer than a yard. **Mass*** is a measure of the amount of material in an object. The SI base unit of mass is the kilogram (kg), which is equal to about 2.2 pounds (lb). This base unit is unusual because it uses a prefix, *kilo-*, instead of the word *gram* alone. We obtain other units for mass by adding prefixes to the word *gram*.

SAMPLE EXERCISE 1.2 Using SI Prefixes

What is the name of the unit that equals **(a)** 10^{-9} gram, **(b)** 10^{-6} second, **(c)** 10^{-3} meter?

SOLUTION

We can find the prefix related to each power of ten in Table 1.5: **(a)** nanogram, ng; **(b)** microsecond, μs; **(c)** millimeter, mm.

Practice Exercise 1

Which of the following weights would you expect to be suitable for weighing on an ordinary bathroom scale?
(a) 2.0×10^7 mg, **(b)** 2500 μg, **(c)** 5×10^{-4} kg, **(d)** 4×10^6 cg, **(e)** 5.5×10^8 dg.

Practice Exercise 2

(a) How many picometers are there in 1 m? **(b)** Express 6.0×10^3 m using a prefix to replace the power of ten. **(c)** Use exponential notation to express 4.22 mg in grams. **(d)** Use decimal notation to express 4.22 mg in grams.

Temperature

Temperature, a measure of the hotness or coldness of an object, is a physical property that determines the direction of heat flow. Heat always flows spontaneously from a substance at higher temperature to one at lower temperature. Thus, the influx of heat we feel when we touch a hot object tells us that the object is at a higher temperature than our hand.

The temperature scales commonly employed in science are the Celsius and Kelvin scales. The **Celsius scale** was originally based on the assignment of 0 °C to the freezing point of water and 100 °C to its boiling point at sea level (**Figure 1.17**).

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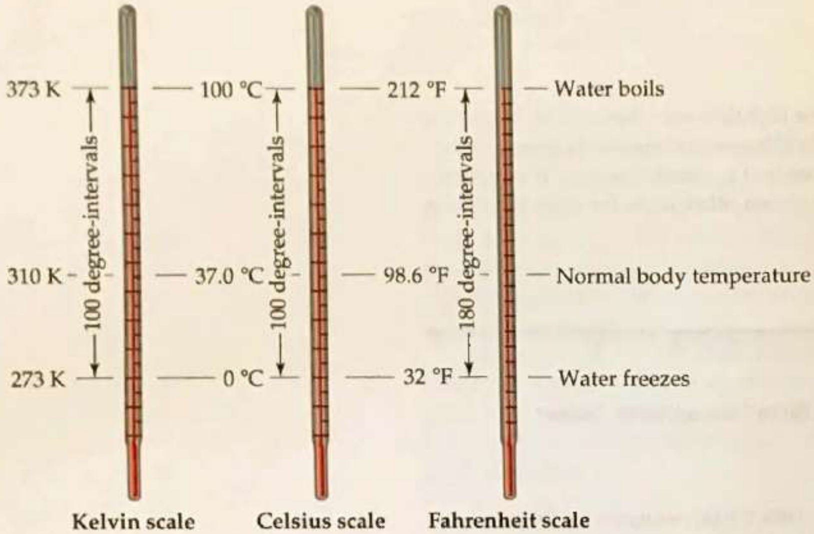
*Mass and weight are often incorrectly thought to be the same. The weight of an object is the force that is exerted on its mass by gravity. In space, where gravitational forces are very weak, an astronaut can be weightless, but he or she cannot be massless. The astronaut's mass in space is the same as it is on Earth.

GO FIGURE

True or false: The "size" of a degree on the Celsius scale is the same as the "size" of a degree on the Kelvin scale.



▲ **Figure 1.17 Comparison of the Kelvin, Celsius, and Fahrenheit temperature scales.**

The **Kelvin scale** is the SI temperature scale, and the SI unit of temperature is the *kelvin* (K). Zero on the Kelvin scale is the lowest attainable temperature, referred to as **absolute zero**. On the Celsius scale, absolute zero has the value, −273.15 °C. The Celsius and Kelvin scales have equal-sized units—that is, a kelvin is the same size as a degree Celsius. Thus, the Kelvin and Celsius scales are related according to

$$K = {}^\circ C + 273.15 \qquad [1.1]$$

The freezing point of water, 0 °C, is 273.15 K (Figure 1.17). Notice that we do not use a degree sign (°) with temperatures on the Kelvin scale.

The common temperature scale in the United States is the *Fahrenheit scale*, which is not generally used in science. Water freezes at 32 °F and boils at 212 °F. The Fahrenheit and Celsius scales are related according to

$$^\circ C = \frac{5}{9}(^\circ F - 32) \quad \text{or} \quad ^\circ F = \frac{9}{5}(^\circ C) + 32 \qquad [1.2]$$

SAMPLE EXERCISE 1.3 Converting Units of Temperature

A weather forecaster predicts the temperature will reach 31 °C. What is this temperature **(a)** in K, **(b)** in °F?

SOLUTION

(a) Using Equation 1.1, we have K = 31 + 273 = 304 K.

(b) Using Equation 1.2, we have

$$^\circ F = \frac{9}{5}(31) + 32 = 56 + 32 = 88\ ^\circ F.$$

Practice Exercise 1

Using Wolfram Alpha (http://www.wolframalpha.com/) or some other reference, determine which of these elements would be liquid at 525 K (assume samples are protected from air): **(a)** bismuth, Bi; **(b)** platinum, Pt; **(c)** selenium, Se; **(d)** calcium, Ca; **(e)** copper, Cu.

Practice Exercise 2

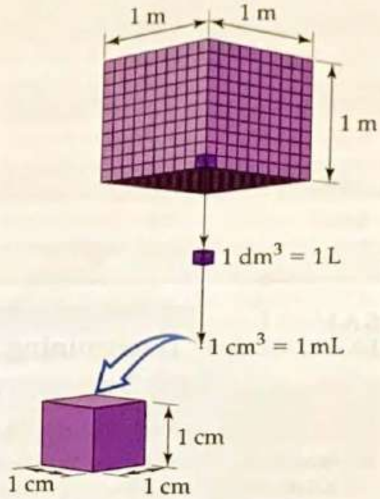
Ethylene glycol, the major ingredient in antifreeze, freezes at −11.5 °C. What is the freezing point in **(a)** K, **(b)** °F?

Derived SI Units

The SI base units are used to formulate *derived units*. A **derived unit** is obtained by multiplication or division of one or more of the base units. We begin with the defining equation for a quantity and, then substitute the appropriate base units. For example, speed is defined as the ratio of distance traveled to elapsed time. Thus, the SI unit for speed—m/s, read "meters per second"—is a derived unit, the SI unit for distance (length), m, divided by the SI unit for time, s. Two common derived units in chemistry are those for volume and density.

GO FIGURE

How many 1-L bottles are required to contain 1 m^3 of liquid?



▲ **Figure 1.18 Volume relationships.** The volume occupied by a cube 1 m on each edge is one cubic meter, 1 m^3. Each cubic meter contains 1000 dm^3. One liter is the same volume as one cubic decimeter, $1\ L = 1\ dm^3$. Each cubic decimeter contains 1000 cubic centimeters, $1\ dm^3 = 1000\ cm^3$. One cubic centimeter equals one milliliter, $1\ cm^3 = 1\ mL$.

Volume

The *volume* of a cube is its length cubed, length^3. Thus, the derived SI unit of volume is the SI unit of length, m, raised to the third power. The cubic meter, m^3, is the volume of a cube that is 1 m on each edge (► **Figure 1.18**). Smaller units, such as cubic centimeters, cm^3 (sometimes written cc), are frequently used in chemistry. Another volume unit used in chemistry is the *liter* (L), which equals a cubic decimeter, dm^3, and is slightly larger than a quart. (The liter is the first metric unit we have encountered that is *not* an SI unit.) There are 1000 milliliters (mL) in a liter, and 1 mL is the same volume as 1 cm^3: $1\ mL = 1\ cm^3$. The devices used most frequently in chemistry to measure volume are illustrated in ▼ **Figure 1.19**.

Syringes, burettes, and pipettes deliver amounts of liquids with more precision than graduated cylinders. Volumetric flasks are used to contain specific volumes of liquid.

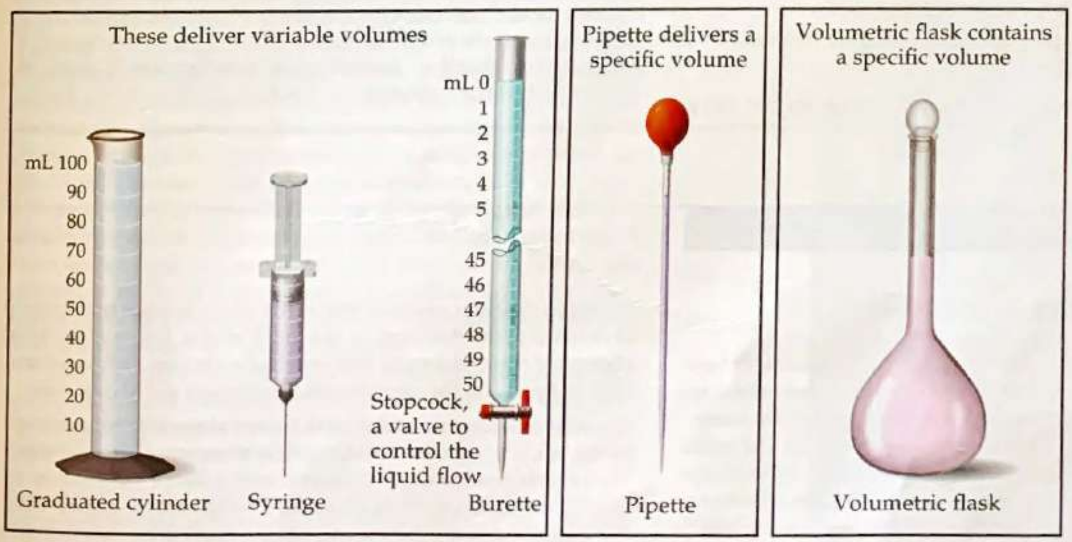
Give It Some Thought

Which of the following quantities represents volume measurement: $15\ m^2$; $2.5 \times 10^2\ m^3$; 5.77 L/s? How do you know?

Density

Density is defined as the amount of mass in a unit volume of a substance:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \qquad [1.3]$$



▲ **Figure 1.19 Common volumetric glassware.**

Table 1.6 **Densities of Selected Substances at 25 °C**

Substance	Density (g/cm^3)
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

The densities of solids and liquids are commonly expressed in either grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). The densities of some common substances are listed in ◀ **Table 1.6**. It is no coincidence that the density of water is 1.00 g/mL; the gram was originally defined as the mass of 1 mL of water at a specific temperature. Because most substances change volume when they are heated or cooled, densities are temperature dependent, and so temperature should be specified when reporting densities. If no temperature is reported, we assume 25 °C, close to normal room temperature.

The terms *density* and *weight* are sometimes confused. A person who says that iron weighs more than air generally means that iron has a higher density than air—1 kg of air has the same mass as 1 kg of iron, but the iron occupies a smaller volume, thereby giving it a higher density. If we combine two liquids that do not mix, the less dense liquid will float on the denser liquid.

SAMPLE EXERCISE 1.4 Determining Density and Using Density to Determine Volume or Mass

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(b) Calculate the volume of 65.0 g of liquid methanol (wood alcohol) if its density is 0.791 g/mL.

(c) What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) if the length of the cube is 2.00 cm?

SOLUTION

(a) We are given mass and volume, so Equation 1.3 yields

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.00 \times 10^2\ \text{g}}{7.36\ \text{cm}^3} = 13.6\ \text{g/cm}^3$$

(b) Solving Equation 1.3 for volume and then using the given mass and density gives $\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{65.0\ \text{g}}{0.791\ \text{g/mL}} = 82.2\ \text{mL}$

(c) We can calculate the mass from the volume of the cube and its density. The volume of a cube is given by its length cubed:

$$\text{Volume} = (2.00\ \text{cm})^3 = (2.00)^3\ \text{cm}^3 = 8.00\ \text{cm}^3$$

Solving Equation 1.3 for mass and substituting the volume and density of the cube, we have

$$\text{Mass} = \text{volume} \times \text{density} = (8.00\ \text{cm}^3)(19.32\ \text{g/cm}^3) = 155\ \text{g}$$

Practice Exercise 1

Platinum, Pt, is one of the rarest of the metals. Worldwide annual production is only about 130 tons. **(a)** Platinum has a density of 21.4 g/cm^3. If thieves were to steal platinum from a bank using a small truck with a maximum payload of 900 lb, how many 1 L bars of the metal could they make off with? **(a)** 19 bars, **(b)** 2 bars, **(c)** 42 bars, **(d)** 1 bar, **(e)** 47 bars.

Practice Exercise 2

(a) Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm^3. **(b)** A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g/mL, how many milliliters of ethanol are needed? **(c)** What is the mass, in grams, of 25.0 mL of mercury (density = 13.6 g/mL)?

Chemistry Put to Work

Chemistry in the News

Because chemistry is so central to our lives, reports on matters of chemical significance appear in the news nearly every day. Some reports tell of breakthroughs in the development of new pharmaceuticals, materials, and processes. Others deal with energy, environmental, and public safety issues. As you study chemistry, you will develop the skills to better understand the importance of chemistry in your life. Here are summaries of a few recent stories in which chemistry plays an important role.

Clean energy from fuel cells. In fuel cells, the energy of a chemical reaction is converted directly into electrical energy. Although fuel cells have long been known as potentially valuable sources of electrical energy, their costs have kept them from widespread use. However, recent advances in technology have brought fuel cells to the fore as sources of reliable and clean electrical power in certain critical situations. They are especially valuable in powering data centers which consume large amounts of electrical power that must be absolutely reliable. For example, failure of electrical power at a major data center for a company such as Amazon, eBay, or Apple could be calamitous for the company and its customers.

eBay recently contracted to build the next phase of its major data center in Utah, utilizing solid–state fuel cells as the source of electrical power. The fuel cells, manufactured by Bloom Energy, a Silicon Valley startup, are large industrial devices about the size of a refrigerator (▶ **Figure 1.20**). The eBay installation utilizes biogas, which consists of methane and other fuel gases derived from landfills and farms. The fuel is combined with oxygen, and the mixture run through a special solid–state device to produce electricity. Because the electricity is being produced close to the data center, transmission of the electrical power from source to consumption is more efficient. In contrast to electrical backup systems employed in the past, the new power source will be the *primary* source of power, operating

GO FIGURE

How would the darts be positioned on the target for the case of "good accuracy, poor precision"?

Good accuracy
Good precision

Poor accuracy
Good precision

Poor accuracy
Poor precision

▲ **Figure 1.22 Precision and accuracy.**

High precision can be achieved on a scale like this one, which has 0.1 milligram accuracy.

1.5 | Uncertainty in Measurement

Two kinds of numbers are encountered in scientific work: *exact numbers* (those whose values are known exactly) and *inexact numbers* (those whose values have some uncertainty). Most of the exact numbers we will encounter in this book have defined values. For example, there are exactly 12 eggs in a dozen, exactly 1000 g in a kilogram, and exactly 2.54 cm in an inch. The number 1 in any conversion factor, such as 1 m = 100 cm or 1 kg = 2.2046 lb, is an exact number. Exact numbers can also result from counting objects. For example, we can count the exact number of marbles in a jar or the exact number of people in a classroom.

Numbers obtained by measurement are always *inexact*. The equipment used to measure quantities always has inherent limitations (equipment errors), and there are differences in how different people make the same measurement (human errors). Suppose ten students with ten balances are to determine the mass of the same dime. The ten measurements will probably vary slightly for various reasons. The balances might be calibrated slightly differently, and there might be differences in how each student reads the mass from the balance. Remember: *Uncertainties always exist in measured quantities.*

Give It Some Thought

Which of the following is an inexact quantity?
(a) the number of people in your chemistry class
(b) the mass of a penny
(c) the number of grams in a kilogram

Precision and Accuracy

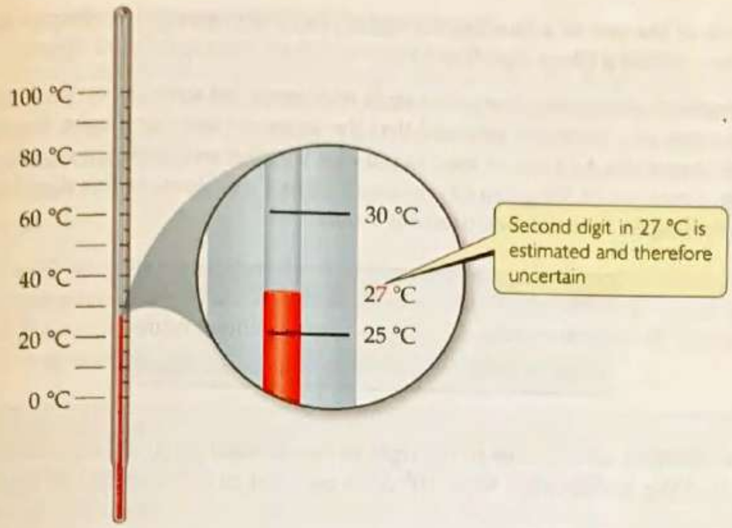
The terms *precision* and *accuracy* are often used in discussing the uncertainties of measured values. **Precision** is a measure of how closely individual measurements agree with one another. **Accuracy** refers to how closely individual measurements agree with the correct, or "true," value. The dart analogy in ◀ **Figure 1.22** illustrates the difference between these two concepts.

In the laboratory we often perform several "trials" of an experiment and average the results. The precision of the measurements is often expressed in terms of the *standard deviation* (Appendix A.5), which reflects how much the individual measurements differ from the average. We gain confidence in our measurements if we obtain nearly the same value each time—that is, when the standard deviation is small. Figure 1.22 reminds us, however, that precise measurements can be inaccurate. For example, if a very sensitive balance is poorly calibrated, the masses we measure will be consistently either high or low. They will be inaccurate even if they are precise.

Significant Figures

Suppose you determine the mass of a dime on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as 2.2405 $\pm$ 0.0001 g. The $\pm$ notation (read "plus or minus") expresses the magnitude of the uncertainty of your measurement. In much scientific work we drop the $\pm$ notation with the understanding that *there is always some uncertainty in the last digit reported for any measured quantity.*

▶ **Figure 1.23** shows a thermometer with its liquid column between two scale marks. We can read the certain digits from the scale and estimate the uncertain one. Seeing that the liquid is between the 25° and 30 °C marks, we estimate the temperature to be 27 °C, being uncertain of the second digit of our measurement. By *uncertain* we mean that the temperature is reliably 27 °C and not 28° or 26 °C, but we can't say that it is *exactly* 27 °C.



◀ **Figure 1.23** **Uncertainty and significant figures in a measurement.**

All digits of a measured quantity, including the uncertain one, are called **significant figures**. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures. The greater the number of significant figures, the greater the precision implied for the measurement.

SAMPLE EXERCISE 1.5 Relating Significant Figures to the Uncertainty of a Measurement

What difference exists between the measured values 4.0 and 4.00 g?

SOLUTION

The value 4.0 has two significant figures, whereas 4.00 has three. This difference implies that 4.0 has more uncertainty. A mass reported as 4.0 g indicates that the uncertainty is in the first decimal place. Thus, the mass is closer to 4.0 than to 3.9 or 4.1 g. We can represent this uncertainty by writing the mass as 4.0 ± 0.1 g. A mass reported as 4.00 g indicates that the uncertainty is in the second decimal place. In this case the mass is closer to 4.00 than 3.99 or 4.01 g, and we can represent it as 4.00 ± 0.01 g. (Without further information, we cannot be sure whether the difference in uncertainties of the two measurements reflects the precision or the accuracy of the measurement.)

Practice Exercise 1

Mo Farah won the 10,000 meter race in the 2012 Olympics with an official time of 27 minutes, 30.42 s. To the correct number of significant figures, what was Farah's average speed in m/sec?
(a) 0. 6059 m/s, **(b)** 1.65042 m/s, **(c)** 6.059064 m/s, **(d)** 0.165042 m/s, **(e)** 6.626192 m/s.

Practice Exercise 2

A sample that has a mass of about 25 g is weighed on a balance that has a precision of ± 0.001 g. How many significant figures should be reported for this measurement?

Give It Some Thought

A digital bathroom scale gives you the following four readings in a row: 155.2, 154.8, 154.9, 154.8 lbs. How would you record your weight?

To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero. *In any measurement that is properly reported, all nonzero digits are significant.* Because zeros can be used either as part of the measured value or merely to locate the decimal point, they may or may not be significant:

1. Zeros *between* nonzero digits are always significant—1005 kg (four significant figures); 7.03 cm (three significant figures).
2. Zeros *at the beginning* of a number are never significant; they merely indicate the position of the decimal point—0.02 g (one significant figure); 0.0026 cm (two significant figures).

3. Zeros *at the end* of a number are significant if the number contains a decimal point—0.0200 g (three significant figures); 3.0 cm (two significant figures).

A problem arises when a number ends with zeros but contains no decimal point. In such cases, it is normally assumed that the zeros are not significant. Exponential notation (Appendix A.1) can be used to indicate whether end zeros are significant. For example, a mass of 10,300 g can be written to show three, four, or five significant figures depending on how the measurement is obtained:

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

In these numbers all the zeros to the right of the decimal point are significant (rules 1 and 3). (The exponential term 10^4 does not add to the number of significant figures.)

SAMPLE EXERCISE 1.6 Assigning Appropriate Significant Figures

The state of Colorado is listed in a road atlas as having a population of 4,301,261 and an area of 104,091 square miles. Do the numbers of significant figures in these two quantities seem reasonable? If not, what seems to be wrong with them?

SOLUTION

The population of Colorado must vary from day to day as people move in or out, are born, or die. Thus, the reported number suggests a much higher degree of *accuracy* than is possible. Secondly, it would not be feasible to actually count every individual resident in the state at any given time. Thus, the reported number suggests far greater *precision* than is possible. A reported number of 4,300,000 would better reflect the actual state of knowledge.

The area of Colorado does not normally vary from time to time, so the question here is whether the accuracy of the measurements is good to six significant figures. It would be possible to achieve such accuracy using satellite technology, provided the legal boundaries are known with sufficient accuracy.

Practice Exercise 1

Which of the following numbers in your personal life are exact numbers?
(a) Your cell phone number, **(b)** your weight, **(c)** your IQ, **(d)** your driver's license number, **(e)** the distance you walked yesterday.

Practice Exercise 2

The back inside cover of the book tells us that there are 5280 ft in 1 mile. Does this make the mile an exact distance?

SAMPLE EXERCISE 1.7 Determining the Number of Significant Figures in a Measurement

How many significant figures are in each of the following numbers (assume that each number is a measured quantity)? **(a)** 4.003, **(b)** 6.023×10^{23}, **(c)** 5000.

SOLUTION

(a) Four; the zeros are significant figures. **(b)** Four; the exponential term does not add to the number of significant figures. **(c)** One; we assume that the zeros are not significant when there is no decimal point shown. If the number has more significant figures, a decimal point should be employed or the number written in exponential notation. Thus, 5000. has four significant figures, whereas 5.00×10^3 has three.

Practice Exercise 1

Sylvia feels as though she may have a fever. Her normal body temperature is 98.7 °F. She measures her body temperature with a thermometer placed under her tongue and gets a value of 102.8 °F. How many significant figures are in this measurement?
(a) Three, the number of degrees to the left of the decimal point; **(b)** four, the number of digits in the measured reading; **(c)** two, the number of digits in the difference between her current reading and her normal body temperature; **(d)** three, the number of digits in her normal body temperature; **(e)** one, the number of digits to the right of the decimal point in the measured value.

Practice Exercise 2

How many significant figures are in each of the following measurements? **(a)** 3.549 g, **(b)** 2.3×10^4 cm, **(c)** 0.00134 m^3.

Significant Figures in Calculations

When carrying measured quantities through calculations, *the least certain measurement limits the certainty of the calculated quantity and thereby determines the number of significant figures in the final answer.* The final answer should be reported with only one uncertain digit. To keep track of significant figures in calculations, we will make frequent use of two rules: one for addition and subtraction, and another for multiplication and division.

1. ***For addition and subtraction,* the result has the same number of decimal places as the measurement with the fewest decimal places.** When the result contains more than the correct number of significant figures, it must be rounded off. Consider the following example in which the uncertain digits appear in color:

This number limits	20.42	⟵ two decimal places
the number of significant	1.322	⟵ three decimal places
figures in the result ⟶	83.1	⟵ one decimal place
	104.842	⟵ round off to one decimal place (104.8)

We report the result as 104.8 because 83.1 has only one decimal place.

2. ***For multiplication and division,* the result contains the same number of significant figures as the measurement with the fewest significant figures.** When the result contains more than the correct number of significant figures, it must be rounded off. For example, the area of a rectangle whose measured edge lengths are 6.221 and 5.2 cm should be reported with two significant figures, 32 cm^2, even though a calculator shows the product to have more digits:

$$\text{Area} = (6.221\text{ cm})(5.2\text{ cm}) = 32.3492\text{ cm}^2 \Rightarrow \text{round off to } 32\text{ cm}^2$$

because 5.2 has two significant figures.

Notice that for addition and subtraction, decimal places are counted in determining how many digits to report in an answer, whereas for multiplication and division, significant figures are counted in determining how many digits to report.

In determining the final answer for a calculated quantity, *exact numbers* are assumed to have an infinite number of significant figures. Thus, when we say, "There are 12 inches in 1 foot," the number 12 is exact, and we need not worry about the number of significant figures in it.

In *rounding off numbers*, look at the leftmost digit to be removed:

- If the leftmost digit removed is less than 5, the preceding number is left unchanged. Thus, rounding off 7.248 to two significant figures gives 7.2.
- If the leftmost digit removed is 5 or greater, the preceding number is increased by 1. Rounding off 4.735 to three significant figures gives 4.74, and rounding 2.376 to two significant figures gives 2.4.*

Give It Some Thought

A rectangular garden plot is measured to be 25.8 m by 18 m. Which of these dimensions needs to be measured to greater accuracy to provide a more accurate estimate of the area of the plot?

*Your instructor may want you to use a slight variation on the rule when the leftmost digit to be removed is exactly 5, with no following digits or only zeros following. One common practice is to round up to the next higher number if that number will be even and down to the next lower number otherwise. Thus, 4.7350 would be rounded to 4.74, and 4.7450 would also be rounded to 4.74.

SAMPLE EXERCISE 1.8 Determining the Number of Significant Figures in a Calculated Quantity

The width, length, and height of a small box are 15.5, 27.3, and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

SOLUTION

In reporting the volume, we can show only as many significant figures as given in the dimension with the fewest significant figures, which is that for the height (two significant figures):

$$\begin{aligned}\text{Volume} &= \text{width} \times \text{length} \times \text{height} \\ &= (15.5\text{ cm})(27.3\text{ cm})(5.4\text{ cm}) \\ &= 2285.01\text{ cm}^3 \Rightarrow 2.3 \times 10^3\text{ cm}^3\end{aligned}$$

A calculator used for this calculation shows 2285.01, which we must round off to two significant figures. Because the resulting number is 2300, it is best reported in exponential notation, 2.3×10^3, to clearly indicate two significant figures.

Practice Exercise 1

Ellen recently purchased a new hybrid car and wants to check her gas mileage. At an odometer setting of 651.1 mi, she fills the tank. At 1314.4 mi she requires 16.1 gal to refill the tank. Assuming that the tank is filled to the same level both times, how is the gas mileage best expressed? **(a)** 40 mi/gal, **(b)** 41 mi/gal, **(c)** 41.2 mi/gal, **(d)** 41.20 mi/gal.

Practice Exercise 2

It takes 10.5 s for a sprinter to run 100.00 m. Calculate her average speed in meters per second and express the result to the correct number of significant figures.

SAMPLE EXERCISE 1.9 Determining the Number of Significant Figures in a Calculated Quantity

A vessel containing a gas at 25 °C is weighed, emptied, and then reweighed as depicted in ▼ **Figure 1.24**. From the data provided, calculate the density of the gas at 25 °C.

SOLUTION

To calculate the density, we must know both the mass and the volume of the gas. The mass of the gas is just the difference in the masses of the full and empty container:

$$(837.63 - 836.25)\text{ g} = 1.38\text{ g}$$

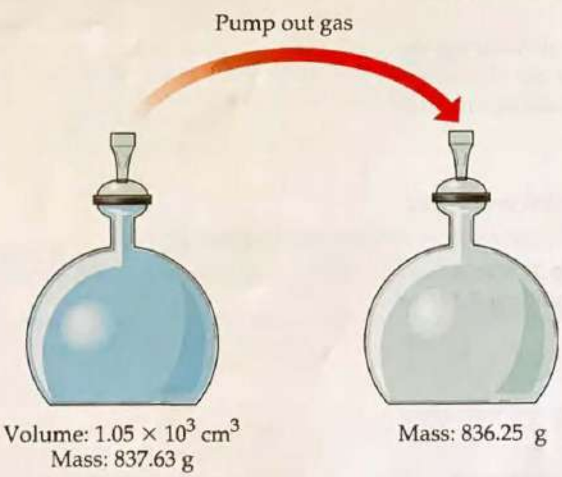
In subtracting numbers, we determine the number of significant figures in our result by counting decimal places in each quantity. In this case each quantity has two decimal places. Thus, the mass of the gas, 1.38 g, has two decimal places.

Using the volume given in the question, $1.05 \times 10^3\text{ cm}^3$, and the definition of density, we have

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.38\text{ g}}{1.05 \times 10^3\text{ cm}^3}$$

$$= 1.31 \times 10^{-3}\text{ g/cm}^3 = 0.00131\text{ g/cm}^3$$

In dividing numbers, we determine the number of significant figures our result should contain by counting the number of significant figures in each quantity. There are three significant figures in our answer, corresponding to the number of significant figures in the two numbers that form the ratio. Notice that in this example, following the rules for determining significant figures gives an answer containing only three significant figures, even though the measured masses contain five significant figures.



▲ **Figure 1.24 Uncertainty and significant figures in a measurement.**

Practice Exercise 1

Which of the following numbers is correctly rounded to three significant figures, as shown in brackets? **(a)** 12,556 [12,500], **(b)** 4.5671×10^{-9} [4.567×10^{-9}], **(c)** 3.00072 [3.001], **(d)** 0.006739 [0.00674], **(e)** 5.4589×10^5 [5.459×10^5].

Practice Exercise 2

If the mass of the container in the sample exercise (Figure 1.24) were measured to three decimal places before and after pumping out the gas, could the density of the gas then be calculated to four significant figures?

When a calculation involves two or more steps and you write answers for intermediate steps, retain at least one nonsignificant digit for the intermediate answers. This procedure ensures that small errors from rounding at each step do not combine to affect the final result. When using a calculator, you may enter the numbers one after another, rounding only the final answer. Accumulated rounding-off errors may account for small differences among results you obtain and answers given in the text for numerical problems.

1.6 | Dimensional Analysis

Because measured quantities have units associated with them, it is important to keep track of units as well as numerical values when using the quantities in calculations. Throughout the text we use **dimensional analysis** in solving problems. In **dimensional analysis**, units are multiplied together or divided into each other along with the numerical values. Equivalent units cancel each other. Using dimensional analysis helps ensure that solutions to problems yield the proper units. Moreover, it provides a systematic way of solving many numerical problems and of checking solutions for possible errors.

The key to using dimensional analysis is the correct use of *conversion factors* to change one unit into another. A **conversion factor** is a fraction whose numerator and denominator are the same quantity expressed in different units. For example, 2.54 cm and 1 in. are the same length: 2.54 cm = 1 in. This relationship allows us to write two conversion factors:

$$\frac{2.54\text{ cm}}{1\text{ in.}} \quad \text{and} \quad \frac{1\text{ in.}}{2.54\text{ cm}}$$

We use the first factor to convert inches to centimeters. For example, the length in centimeters of an object that is 8.50 in. long is

$$\text{Number of centimeters} = (8.50\ \cancel{\text{in.}})\ \frac{2.54\text{ cm}}{1\ \cancel{\text{in.}}} = 21.6\text{ cm}$$

Desired unit

Given unit

The unit inches in the denominator of the conversion factor cancels the unit inches in the given data (8.50 *inches*), so that the centimeters unit in the numerator of the conversion factor becomes the unit of the final answer. Because the numerator and denominator of a conversion factor are equal, multiplying any quantity by a conversion factor is equivalent to multiplying by the number 1 and so does not change the intrinsic value of the quantity. The length 8.50 in. is the same as the length 21.6 cm.

In general, we begin any conversion by examining the units of the given data and the units we desire. We then ask ourselves what conversion factors we have available to take us from the units of the given quantity to those of the desired one. When we multiply a quantity by a conversion factor, the units multiply and divide as follows:

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

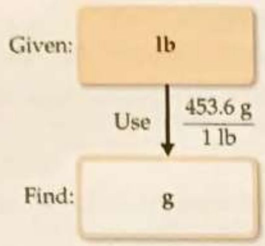
If the desired units are not obtained in a calculation, an error must have been made somewhere. Careful inspection of units often reveals the source of the error.

SAMPLE EXERCISE 1.10 Converting Units

If a woman has a mass of 115 lb, what is her mass in grams? (Use the relationships between units given on the back inside cover of the text.)

SOLUTION

Because we want to change from pounds to grams, we look for a relationship between these units of mass. The conversion factor table found on the back inside cover tells us that 1 lb = 453.6 g.



To cancel pounds and leave grams, we write the conversion factor with grams in the numerator and pounds in the denominator:

$$\text{Mass in grams} = (115\ \text{lb})\left(\frac{453.6\ \text{g}}{1\ \text{lb}}\right) = 5.22 \times 10^4\ \text{g}$$

The answer can be given to only three significant figures, the number of significant figures in 115 lb. The process we have used is diagrammed in the margin.

Practice Exercise 1

At a particular instant in time the Earth is judged to be 92,955,000 miles from the Sun. What is the distance in kilometers to four significant figures? (See back inside cover for conversion factor). **(a)** 5763×10^4 km, **(b)** 1.496×10^8 km, **(c)** 1.49596×10^8 km, **(d)** 1.483×10^4 km, **(e)** 57,759,000 km.

Practice Exercise 2

By using a conversion factor from the back inside cover, determine the length in kilometers of a 500.0-mi automobile race.

Strategies in Chemistry

Estimating Answers

Calculators are wonderful devices; they enable you to get to the wrong answer very quickly. Of course, that's not the destination you want. You can take certain steps to avoid putting that wrong answer into your homework set or on an exam. One is to keep track of the units in a calculation and use the correct conversion factors. Second, you can do a quick mental check to be sure that your answer is reasonable: you can try to make a "ballpark" estimate.

A ballpark estimate involves making a rough calculation using numbers that are rounded off in such a way that the arithmetic can be done without a calculator. Even though this approach does not give an exact answer, it gives one that is roughly the correct size. By using dimensional analysis and by estimating answers, you can readily check the reasonableness of your calculations.

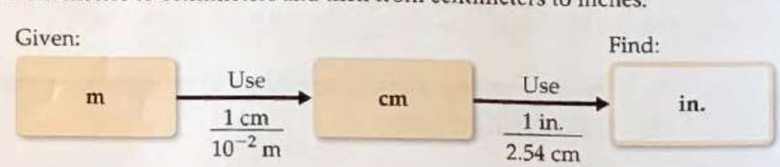
You can get better at making estimates by practicing in everyday life. How far is it from your dorm room to the chemistry lecture hall? How much do your parents pay for gasoline per year? How many bikes are there on campus? If you respond "I have no idea" to these questions, you're giving up too easily. Try estimating familiar quantities and you'll get better at making estimates in science and in other aspects of your life where a misjudgment can be costly.

Give It Some Thought

How do we determine how many digits to use in conversion factors, such as the one between pounds and grams in Sample Exercise 1.10?

Using Two or More Conversion Factors

It is often necessary to use several conversion factors in solving a problem. As an example, let's convert the length of an 8.00-m rod to inches. The table on the back inside cover does not give the relationship between meters and inches. It *does*, however, give the relationship between centimeters and inches (1 in. = 2.54 cm). From our knowledge of SI prefixes, we know that 1 cm = 10^{-2} m. Thus, we can convert step by step, first from meters to centimeters and then from centimeters to inches:



Combining the given quantity (8.00 m) and the two conversion factors, we have

$$\text{Number of inches} = (8.00\ \text{m})\left(\frac{1\ \text{cm}}{10^{-2}\ \text{m}}\right)\left(\frac{1\ \text{in.}}{2.54\ \text{cm}}\right) = 315\ \text{in.}$$

The first conversion factor is used to cancel meters and convert the length to centimeters. Thus, meters are written in the denominator and centimeters in the numerator.

The second conversion factor is used to cancel centimeters and convert the length to inches, so it has centimeters in the denominator and inches, the desired unit, in the numerator.

Note that you could have used 100 cm = 1 m as a conversion factor as well in the second parentheses. As long as you keep track of your given units and cancel them properly to obtain the desired units, you are likely to be successful in your calculations.

SAMPLE EXERCISE 1.11 Converting Units Using Two or More Conversion Factors

The average speed of a nitrogen molecule in air at 25 °C is 515 m/s. Convert this speed to miles per hour.

SOLUTION

To go from the given units, m/s, to the desired units, mi/hr, we must convert meters to miles and seconds to hours. From our knowledge of SI prefixes we know that 1 km = 10^3 m. From the relationships given on the back inside cover of the book, we find that 1 mi = 1.6093 km. Thus, we can convert m to km and then convert km to mi. From our knowledge of time we know that 60 s = 1 min and 60 min = 1 hr. Thus, we can convert s to min and then convert min to hr. The overall process is

Applying first the conversions for distance and then those for time, we can set up one long equation in which unwanted units are canceled:

$$\text{Speed in mi/hr} = \left(515\frac{\cancel{\text{m}}}{\cancel{\text{s}}}\right)\left(\frac{1\ \cancel{\text{km}}}{10^3\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{mi}}{1.6093\ \cancel{\text{km}}}\right)\left(\frac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}}\right)\left(\frac{60\ \cancel{\text{min}}}{1\ \text{hr}}\right)$$

$$= 1.15 \times 10^3\ \text{mi/hr}$$

Our answer has the desired units. We can check our calculation, using the estimating procedure described in the "Strategies in Chemistry" box. The given speed is about 500 m/s. Dividing by 1000 converts m to km, giving 0.5 km/s. Because 1 mi is about 1.6 km, this speed corresponds to 0.5/1.6 = 0.3 mi/s. Multiplying by 60 gives about 0.3 × 60 = 20 mi/min. Multiplying again by 60 gives 20 × 60 = 1200 mi/hr. The approximate solution (about 1200 mi/hr) and the detailed solution (1150 mi/hr) are reasonably close. The answer to the detailed solution has three significant figures, corresponding to the number of significant figures in the given speed in m/s.

Practice Exercise 1

Fabiola, who lives in Mexico City, fills her car with gas, paying 357 pesos for 40.0 L. What is her fuel cost in dollars per gallon, if 1 peso = 0.0759 dollars? **(a)** \$1.18/gal, **(b)** \$3.03/gal, **(c)** \$1.47/gal, **(d)** \$9.68/gal, **(e)** \$2.56/gal.

Practice Exercise 2

A car travels 28 mi per gallon of gasoline. What is the mileage in kilometers per liter?

Conversions Involving Volume

The conversion factors previously noted convert from one unit of a given measure to another unit of the same measure, such as from length to length. We also have conversion factors that convert from one measure to a different one. The density of a substance, for example, can be treated as a conversion factor between mass and volume. Suppose we want to know the mass in grams of 2 cubic inches (2.00 in.^{3}) of gold, which has a density of 19.3 g/cm^3. The density gives us the conversion factors:

$$\frac{19.3\ \text{g}}{1\ \text{cm}^3} \quad \text{and} \quad \frac{1\ \text{cm}^3}{19.3\ \text{g}}$$

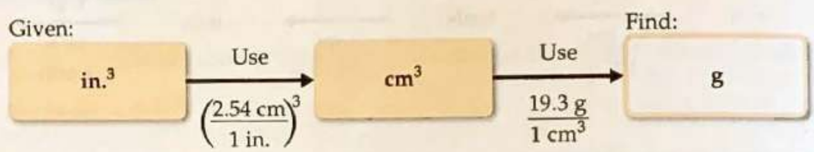
Because we want a mass in grams, we use the first factor, which has mass in grams in the numerator. To use this factor, however, we must first convert cubic inches to cubic centimeters. The relationship between in.^{3} and cm^3 is not given on the back inside cover, but the relationship between inches and centimeters is given: 1 in. = 2.54 cm (exactly). Cubing both sides of this equation gives $(1\text{ in.})^3 = (2.54\text{ cm})^3$, from which we write the desired conversion factor:

$$\frac{(2.54\text{ cm})^3}{(1\text{ in.})^3} = \frac{(2.54)^3\text{ cm}^3}{(1)^3\text{ in.}^3} = \frac{16.39\text{ cm}^3}{1\text{ in.}^3}$$

Notice that both the numbers and the units are cubed. Also, because 2.54 is an exact number, we can retain as many digits of $(2.54)^3$ as we need. We have used four, one more than the number of digits in the density (19.3 g/cm^3). Applying our conversion factors, we can now solve the problem:

$$\text{Mass in grams} = (2.00\text{ in.}^3)\left(\frac{16.39\text{ cm}^3}{1\text{ in.}^3}\right)\left(\frac{19.3\text{ g}}{1\text{ cm}^3}\right) = 633\text{ g}$$

The procedure is diagrammed here. The final answer is reported to three significant figures, the same number of significant figures as in 2.00 in.^{3} and 19.3 g.



SAMPLE EXERCISE 1.12 Converting Volume Units

Earth's oceans contain approximately $1.36 \times 10^9\text{ km}^3$ of water. Calculate the volume in liters.

How many liters of water do Earth's oceans contain?

SOLUTION

From the back inside cover, we find $1\text{ L} = 10^{-3}\text{ m}^3$, but there is no relationship listed involving km^3. From our knowledge of SI prefixes, however, we know $1\text{ km} = 10^3\text{ m}$ and we can use this relationship between lengths to write the desired conversion factor between volumes:

$$\left(\frac{10^3\text{ m}}{1\text{ km}}\right)^3 = \frac{10^9\text{ m}^3}{1\text{ km}^3}$$

Thus, converting from km^3 to m^3 to L, we have

$$\text{Volume in liters} = (1.36 \times 10^9\text{ km}^3)\left(\frac{10^9\text{ m}^3}{1\text{ km}^3}\right)\left(\frac{1\text{ L}}{10^{-3}\text{ m}^3}\right) = 1.36 \times 10^{21}\text{ L}$$

Practice Exercise 1

A barrel of oil as measured on the oil market is equal to 1.333 U.S. barrels. A U.S. barrel is equal to 31.5 gal. If oil is on the market at \$94.0 per barrel, what is the price in dollars per gallon? (a) \$2.24/gal, (b) \$3.98/gal, (c) \$2.98/gal, (d) \$1.05/gal, (e) \$8.42/gal.

Practice Exercise 2

The surface area of Earth is $510 \times 10^6\text{ km}^2$, and 71% of this is ocean. Using the data from the sample exercise, calculate the average depth of the world's oceans in feet.

Strategies in Chemistry

The Importance of Practice

If you have ever played a musical instrument or participated in athletics, you know that the keys to success are practice and discipline. You cannot learn to play a piano merely by listening to music, and you cannot learn how to play basketball merely by watching games on television. Likewise, you cannot learn chemistry by merely watching your instructor give lectures. Simply reading this book, listening to lectures, or reviewing notes will not usually be sufficient when exam time comes around. Your task is to master chemical concepts and practices to a degree that you can put them to use in solving problems and answering questions. Solving problems correctly takes practice—actually, a fair amount of it. You will do well in your chemistry course if you embrace the idea that you need to master the materials presented, and then learn how to apply them in solving problems. Even if you're a brilliant student, this will take time; it's what being a student is all about. Almost no one fully absorbs new material on a first reading, especially when new concepts are being presented. You are sure to more fully master the content of the chapters by reading them through at least twice, even more for passages that present you with difficulties in understanding.

Throughout the book, we have provided sample exercises in which the solutions are shown in detail. For practice exercises, we supply only the answer, at the back of the book. It is important that you use these exercises to test yourself.

The practice exercises in this text and the homework assignments given by your instructor provide the minimal practice that you will need to succeed in your chemistry course. Only by working all the assigned problems will you face the full range of difficulty and coverage that your instructor expects you to master for exams. There is no substitute for a determined and perhaps lengthy effort to work problems on your own. If you are stuck on a problem, however, ask for help from your instructor, a teaching assistant, a tutor, or a fellow student. Spending an inordinate amount of time on a single exercise is rarely effective unless you know that it is particularly challenging and is expected to require extensive thought and effort.

SAMPLE EXERCISE 1.13 Conversions Involving Density

What is the mass in grams of 1.00 gal of water? The density of water is 1.00 g/mL.

SOLUTION

Before we begin solving this exercise, we note the following:

(1) We are given 1.00 gal of water (the known, or given, quantity) and asked to calculate its mass in grams (the unknown).

(2) We have the following conversion factors either given, commonly known, or available on the back inside cover of the text:

$$\frac{1.00 \text{ g water}}{1 \text{ mL water}} \quad \frac{1 \text{ L}}{1000 \text{ mL}} \quad \frac{1 \text{ L}}{1.057 \text{ qt}} \quad \frac{1 \text{ gal}}{4 \text{ qt}}$$

The first of these conversion factors must be used as written (with grams in the numerator) to give the desired result, whereas the last conversion factor must be inverted in order to cancel gallons:

$$\text{Mass in grams} = (1.00 \text{ gal})\left(\frac{4 \text{ qt}}{1 \text{ gal}}\right)\left(\frac{1 \text{ L}}{1.057 \text{ qt}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right)\left(\frac{1.00 \text{ g}}{1 \text{ mL}}\right)$$

$$= 3.78 \times 10^3 \text{ g water}$$

The unit of our final answer is appropriate, and we have taken care of our significant figures. We can further check our calculation by estimating. We can round 1.057 off to 1. Then focusing on the numbers that do not equal 1 gives $4 \times 1000 = 4000$ g, in agreement with the detailed calculation.

You should also use common sense to assess the reasonableness of your answer. In this case we know that most people can lift a gallon of milk with one hand, although it would be tiring to carry it around all day. Milk is mostly water and will have a density not too different from that of water. Therefore, we might estimate that a gallon of water has mass that is more than 5 lb but less than 50 lb. The mass we have calculated, $3.78 \text{ kg} \times 2.2 \text{ lb/kg} = 8.3 \text{ lb}$, is thus reasonable as an order-of-magnitude estimate.

Practice Exercise 1

Trex is a manufactured substitute for wood compounded from post-consumer plastic and wood. It is frequently used in outdoor decks. Its density is reported as 60 lb/ft^3. What is the density of Trex in kg/L? **(a)** 138 kg/L, **(b)** 0.960 kg/L, **(c)** 259 kg/L, **(d)** 15.8 kg/L, **(e)** 11.5 kg/L.

Practice Exercise 2

The density of the organic compound benzene is 0.879 g/mL. Calculate the mass in grams of 1.00 qt of benzene.

A Trex deck.

Strategies in Chemistry

The Features of This Book

If, like most students, you haven't yet read the part of the Preface to this text entitled TO THE STUDENT, *you should do it now*. In less than two pages of reading you will encounter valuable advice on how to navigate your way through this book and through the course. We're serious! This is advice you can use.

The TO THE STUDENT section describes how text features such as "What's Ahead," Key Terms, Learning Outcomes, and Key Equations will help you remember what you have learned. We describe there also how to take advantage of the text's Web site, where many types of online study tools are available. If you have registered for MasteringChemistry®, you will have access to many helpful animations, tutorials, and additional problems correlated to specific topics and sections of each chapter. An interactive eBook is also available online.

As previously mentioned, working exercises is very important—in fact, essential. You will find a large variety of exercises at the end of each chapter that are designed to test your problem-solving skills in chemistry. Your instructor will very likely assign some of these end-of-chapter exercises as homework. The first few exercises called "Visualizing Concepts" are meant to test how well you understand a concept without plugging a lot of numbers into a formula. The other exercises are grouped in pairs, with the answers given at the back of the book to the odd-numbered exercises (those with red exercise numbers). An exercise with a [bracket] around its number is designed to be more challenging. Additional Exercises appear after the regular exercises; the chapter sections that they cover are not identified, and they are not paired. Integrative Exercises, which start appearing from Chapter 3, are problems that require skills learned in previous chapters. Also first appearing in Chapter 3, are Design an Experiment exercises consisting of problem scenarios that challenge you to design experiments to test hypotheses.

Many chemical databases are available, usually on the Web. The *CRC Handbook of Chemistry and Physics* is the standard reference for many types of data and is available in libraries. The *Merck Index* is a standard reference for the properties of many organic compounds, especially ones of biological interest. WebElements (http://www.webelements.com/) is a good Web site for looking up the properties of the elements. Wolfram Alpha (http://www.wolframalpha.com/) can also be a source of useful information on substances, numerical values, and other data.

Chapter Summary and Key Terms

THE STUDY OF CHEMISTRY (SECTION 1.1) **Chemistry** is the study of the composition, structure, properties, and changes of **matter**. The composition of matter relates to the kinds of **elements** it contains. The structure of matter relates to the ways the **atoms** of these elements are arranged. A **property** is any characteristic that gives a sample of matter its unique identity. A **molecule** is an entity composed of two or more atoms with the atoms attached to one another in a specific way.

CLASSIFICATIONS OF MATTER (SECTION 1.2) Matter exists in three physical states, **gas**, **liquid**, and **solid**, which are known as the **states of matter**. There are two kinds of **pure substances**: **elements** and **compounds**. Each element has a single kind of atom and is represented by a chemical symbol consisting of one or two letters, with the first letter capitalized. Compounds are composed of two or more elements joined chemically. The **law of constant composition**, also called the **law of definite proportions**, states that the elemental composition of a pure compound is always the same. Most matter consists of a mixture of substances. **Mixtures** have variable compositions and can be either homogeneous or heterogeneous; homogeneous mixtures are called **solutions**.

PROPERTIES OF MATTER (SECTION 1.3) Each substance has a unique set of **physical properties** and **chemical properties** that can be used to identify it. During a **physical change**, matter does not change its composition. **Changes of state** are physical changes. In a **chemical change (chemical reaction)** a substance is transformed into a chemically different substance. **Intensive properties** are independent of the amount of matter examined and are used to identify substances. **Extensive properties** relate to the amount of substance present. Differences in physical and chemical properties are used to separate substances.

The **scientific method** is a dynamic process used to answer questions about the physical world. Observations and experiments lead to tentative explanations or **hypotheses**. As a hypothesis is tested and refined, a **theory** may be developed that can predict the results of future observations and experiments. When observations repeatedly lead to the same consistent results, we speak of a **scientific law**, a general rule that summarizes how nature behaves.

UNITS OF MEASUREMENT (SECTION 1.4) Measurements in chemistry are made using the **metric system**. Special emphasis is placed on **SI units**, which are based on the meter, the kilogram, and the second as the basic units of length, **mass**, and time, respectively. SI units use prefixes to indicate fractions or multiples of base units. The SI **temperature** scale is the **Kelvin scale**, although the **Celsius scale** is frequently used as well. **Absolute zero** is the lowest temperature attainable. It has the value 0 K. A **derived unit** is obtained by multiplication or division of SI base units. Derived units are needed for defined quantities such as speed or volume. **Density** is an important defined quantity that equals mass divided by volume.

UNCERTAINTY IN MEASUREMENT (SECTION 1.5) All measured quantities are inexact to some extent. The **precision** of a measurement indicates how closely different measurements of a quantity agree with one another. The **accuracy** of a measurement indicates how well a measurement agrees with the accepted or "true" value. The **significant figures** in a measured quantity include one estimated digit, the last digit of the measurement. The significant figures indicate the extent of the uncertainty of the measurement. Certain rules must be followed so that a calculation involving measured quantities is reported with the appropriate number of significant figures.

DIMENSIONAL ANALYSIS (SECTION 1.6) In the **dimensional analysis** approach to problem solving, we keep track of units as we carry measurements through calculations. The units are multiplied together, divided into each other, or canceled like algebraic quantities. Obtaining the proper units for the final result is an important means of checking the method of calculation. When converting units and when carrying out several other types of problems, **conversion factors** can be used. These factors are ratios constructed from valid relations between equivalent quantities.