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*).

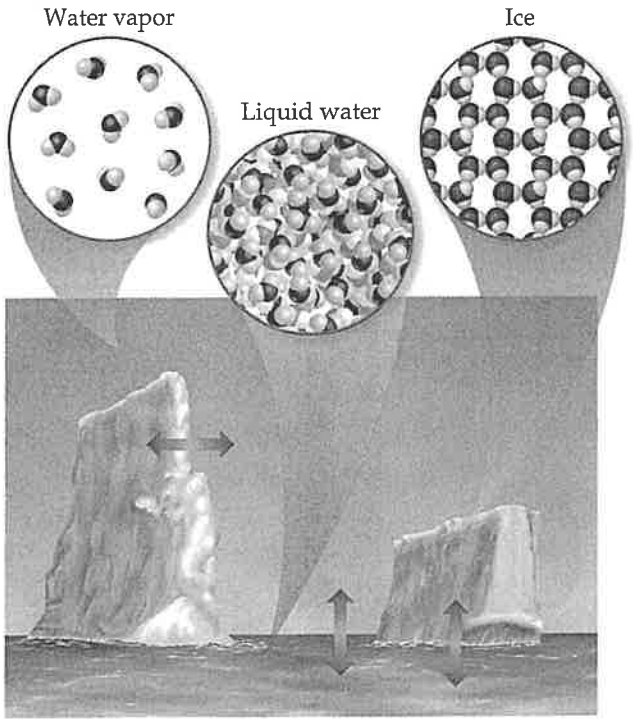
States of Matter

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, assumes the shape of the portion of the container it occupies, and is not compressible to any appreciable extent.
- A **solid** has both a definite shape and a definite volume and is not compressible 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 while the molecules are for the most part locked in place in a solid, they retain considerable freedom of motion in a liquid. 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

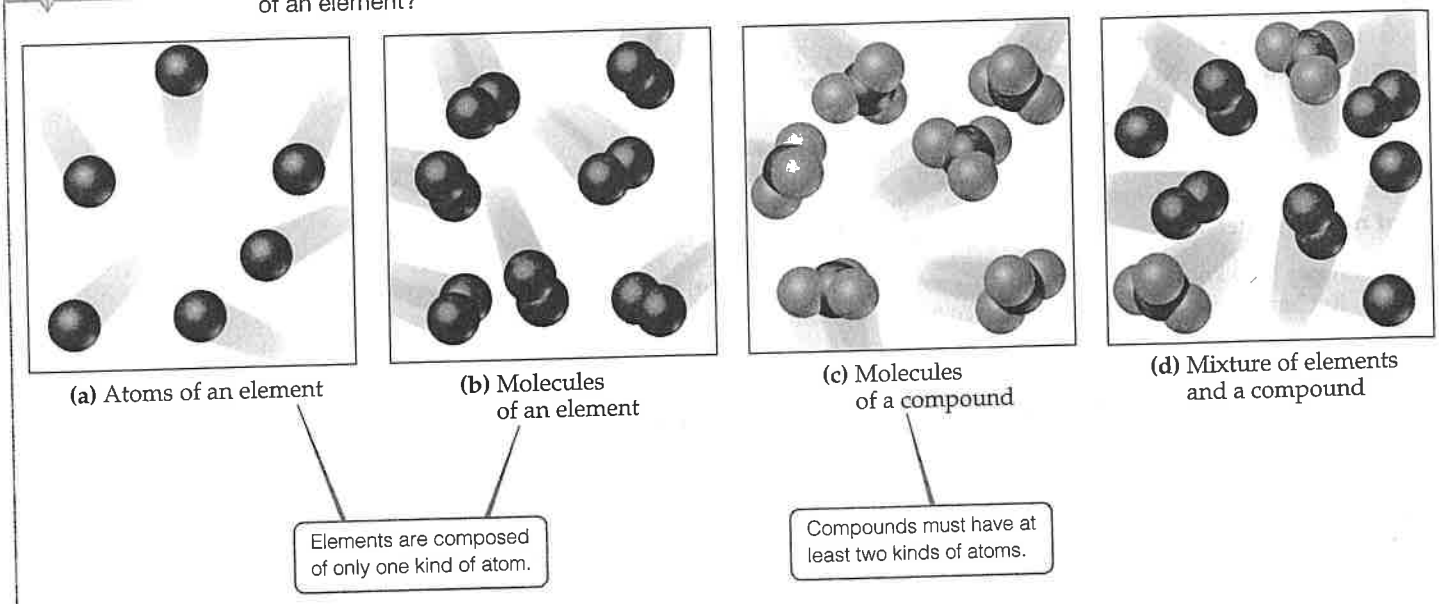
Most forms of matter we encounter—the air we breathe (a gas), the gasoline we burn in our cars (a liquid), and the sidewalk we walk on (a solid)—are not chemically pure. We can, however, separate these forms of matter into pure substances. A **pure substance** (usually referred to simply as a *substance*) 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.

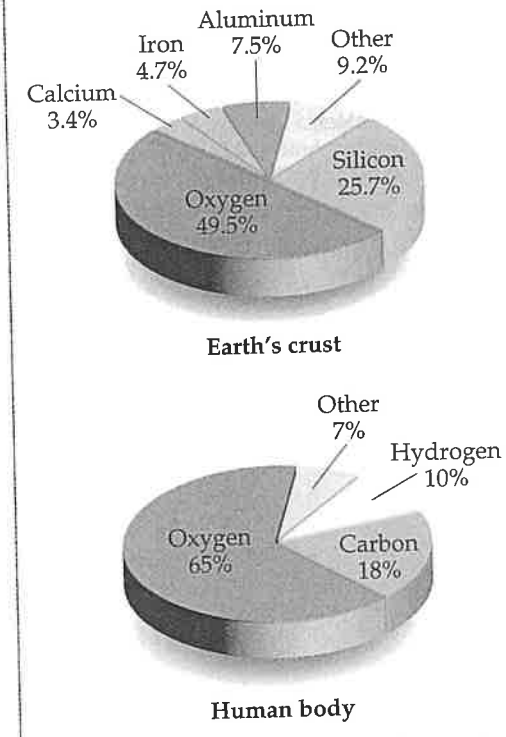
Figure 1.5(**d**) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each substance retains its chemical identity.

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 If the lower pie chart was drawn as the percentage in terms of number of atoms rather than the percentage in terms of mass, would the hydrogen slice of the pie get larger or smaller?



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

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**).

Table 1.1 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 1.1). 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

TABLE 1.1 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*)

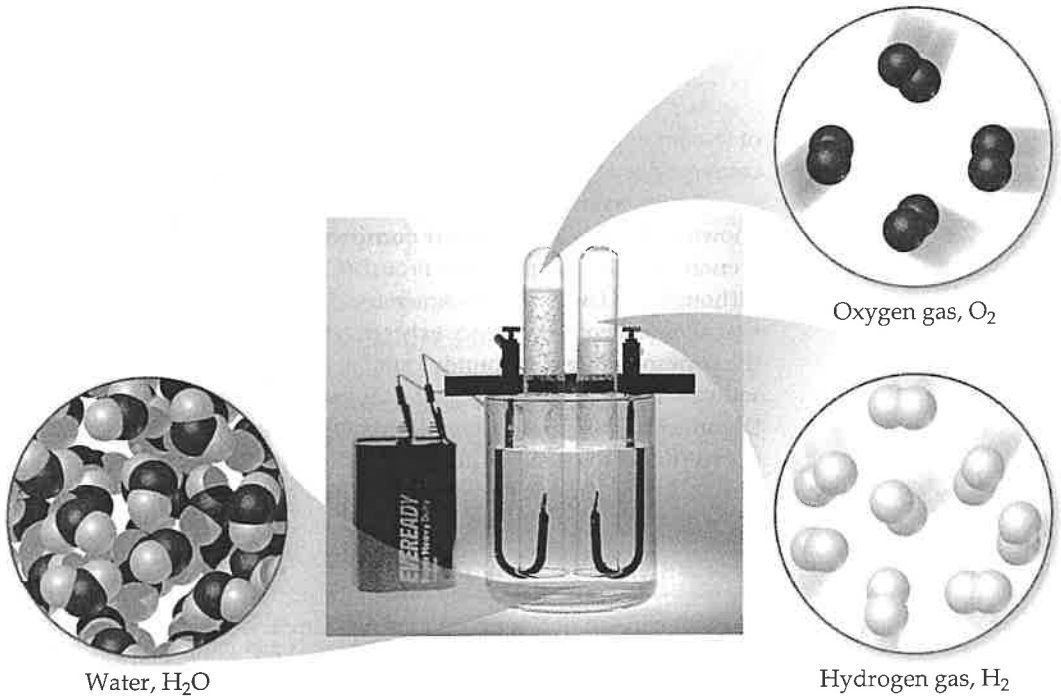
*U.S. Geological Survey Circular 285, U.S Department of the Interior.

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**).

Go Figure Is the volume of H_2 produced larger than the volume of O_2 produced because (a) hydrogen atoms are lighter than oxygen atoms, (b) hydrogen atoms are larger than oxygen atoms, or (c) each water molecule contains one oxygen atom and two hydrogen atoms?



▲ **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.

Decomposing pure water into its constituent elements shows that it contains 11% hydrogen and 89% oxygen by mass, regardless of its source. This ratio is constant because every water molecule has the same number of hydrogen and oxygen atoms. While the mass percentages make it seem that water is mostly oxygen, there are actually two hydrogen atoms and only one oxygen atom per molecule. The explanation for this apparent discrepancy comes from the fact that hydrogen atoms are much lighter than oxygen atoms. This macroscopic composition corresponds to the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom:

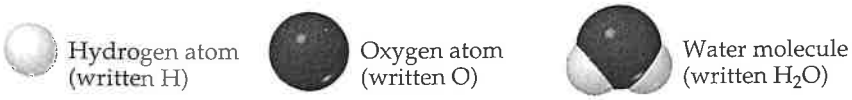


TABLE 1.2 Comparison of Water, Hydrogen, and Oxygen

	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 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 1.2**, 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.

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. This simply is not true. Regardless of its source—nature or a laboratory—a pure compound has the same composition and properties under the same conditions. 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

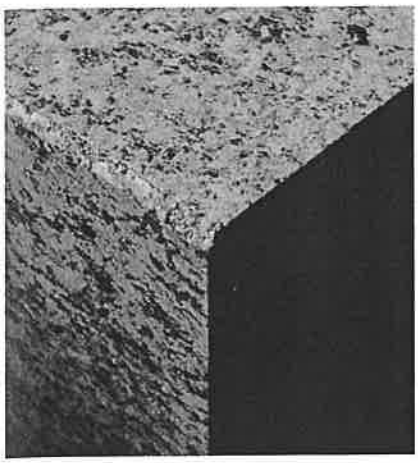
Hydrogen, oxygen, and water are all composed of molecules. What is it about a molecule of water that makes it a compound, whereas hydrogen and oxygen are elements?

Mixtures

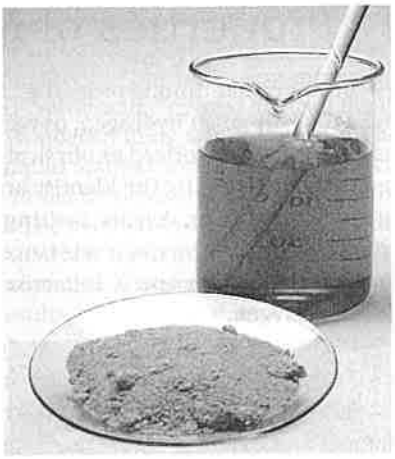
Most of the matter we encounter consists of mixtures of different substances. Each substance in a mixture retains its chemical identity and properties. In contrast to a pure substance, which by definition has a fixed composition, the composition of a mixture can vary. A cup of sweetened coffee, for example, 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.

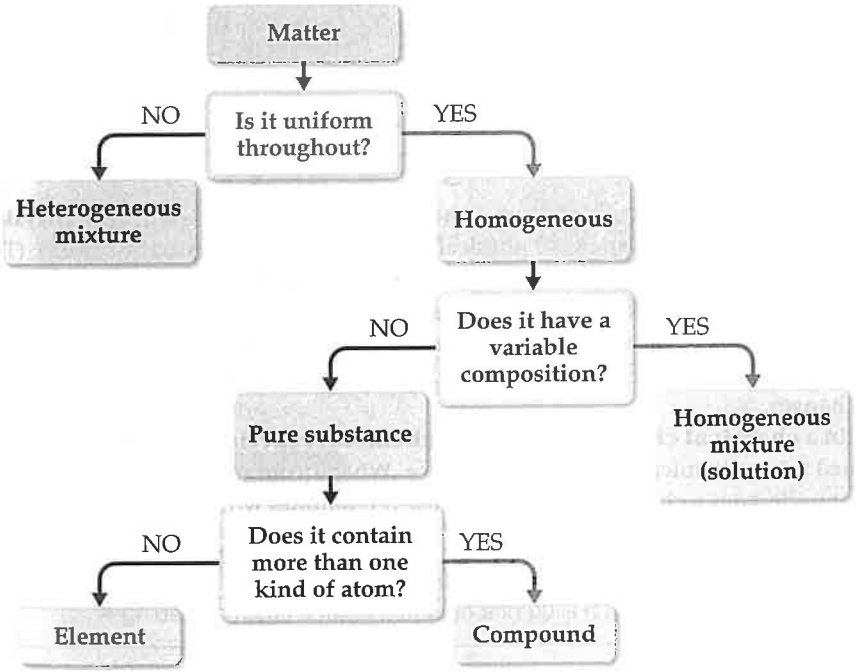


(a)



(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 pentahydrate], 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 the inside of a grapefruit?

(a) It is a pure compound.
(b) It consists of a homogeneous 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 1.2 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

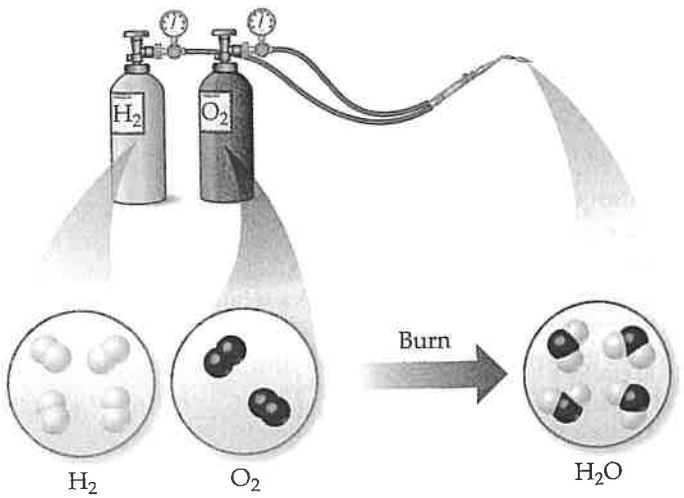
When we say that lead is a denser metal than aluminum, are we talking about an extensive or intensive property?

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.

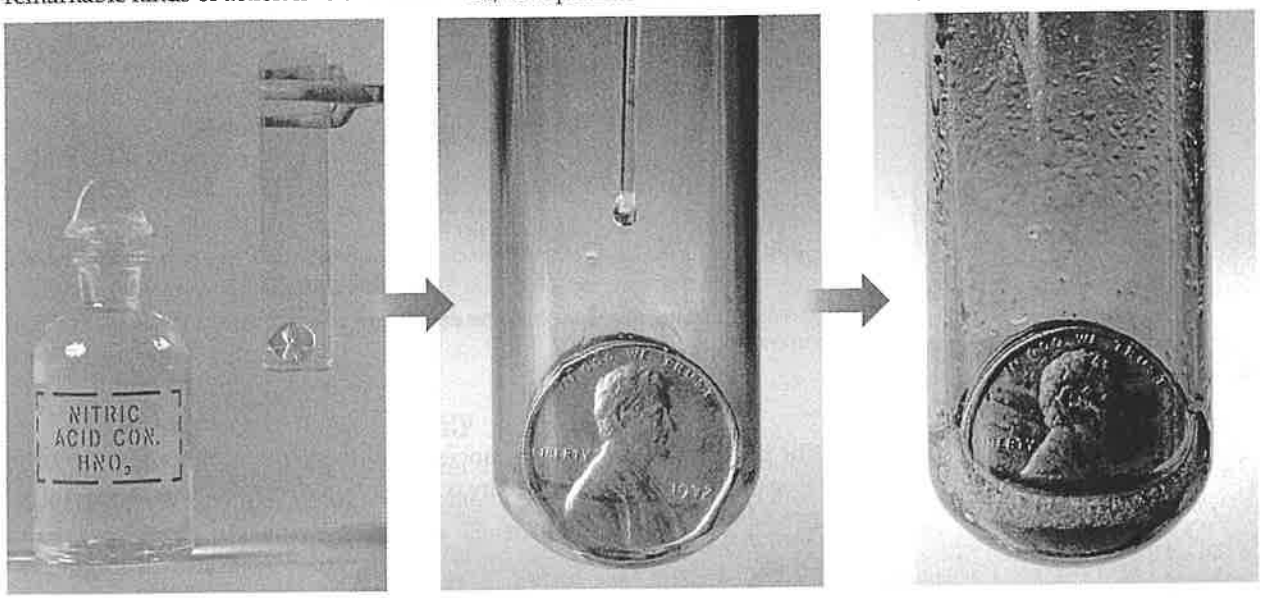
In a **chemical change** (also called a **chemical reaction**), a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water (**Figure 1.10**).

Chemical changes can be dramatic. In the account given in **Figure 1.11**, Ira Remsen, author of a popular chemistry text published in 1901, describes his first experiences with chemical reactions.



▲ Figure 1.10 A chemical reaction.

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.*



▲ **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.

Give It Some Thought

Which of these changes are physical and which are chemical? Explain.

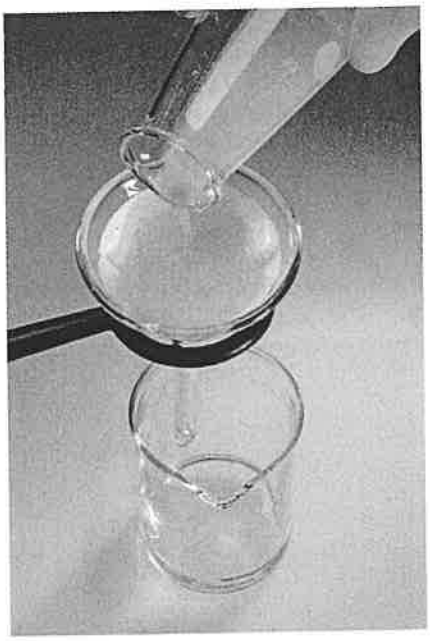
(a) Plants make sugar from carbon dioxide and water.
(b) Water vapor in the air forms frost.
(c) A goldsmith melts a nugget of gold and pulls it into a wire.

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.

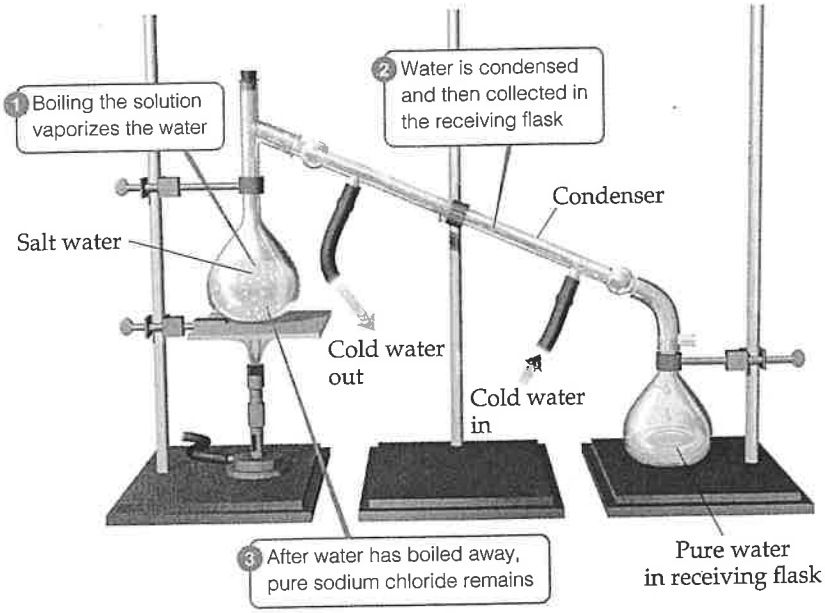
An important method of separating the components of a homogeneous mixture is **distillation**, a process that depends on the different abilities of substances to form gases. For example, if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous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

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





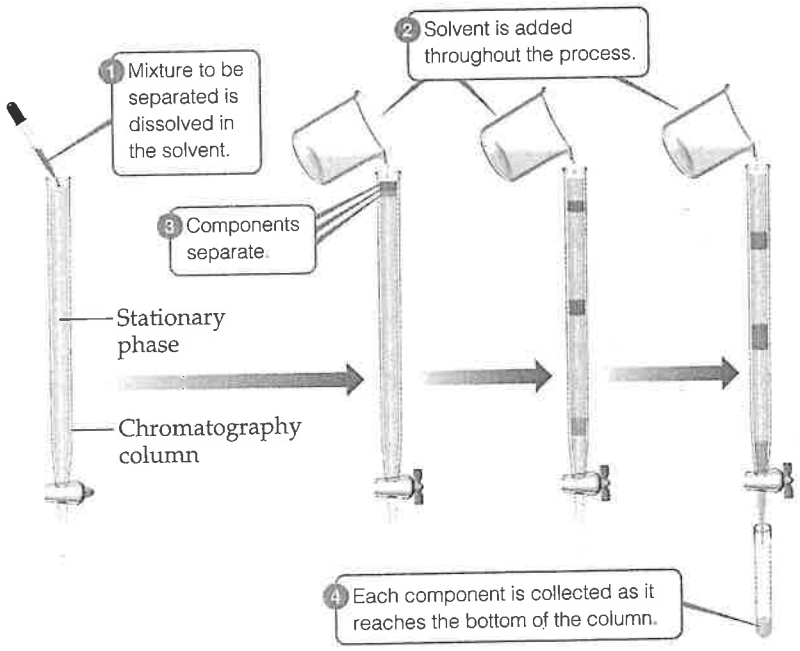
▲ **Figure 1.12 Separation by filtration.** A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.



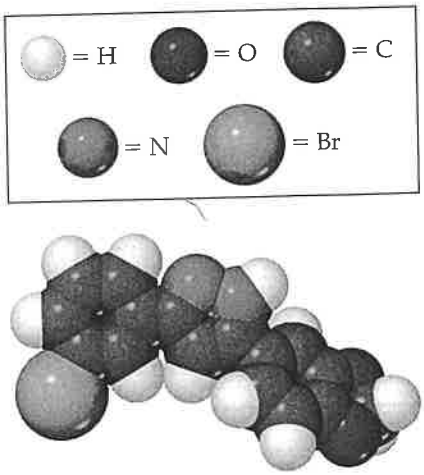
▲ **Figure 1.13 Distillation.** Apparatus for separating a sodium chloride solution (salt water) into its components.

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**.

Go Figure Is the separation of the mixture shown here a physical or chemical process?



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



▲ Figure 1.23 Anle-138b, a promising molecule for the treatment of degenerative brain diseases.

effective therapies. This work illustrates the complex and sometimes tortuous path that takes the scientist from promising beginnings to a much desired goal.

Chemical Approaches to Art Restoration. Works of art are made from materials that over time can undergo chemical reactions that alter their appearance or undermine their mechanical stability. Layers of dirt and pollutants can build up on the surface, discoloring statues, murals, and paintings. Plasters that serve as the base for frescoes react with gases in the atmosphere. One approach to art restoration is to use selected chemical reactions to undo or reverse the effects of detrimental reactions that have occurred over time.

One such example involves murals at the ancient Mayan ruins of Mayapan in Mexico's Yucatán peninsula. In the 1960s, polymer coatings were applied to these murals in an attempt to preserve them, but within a decade it became clear that the polymer coatings were doing more harm than good. Unfortunately, years of oxidation and cross-linking reactions had rendered the polymer coatings insoluble to nearly every type of organic solvent. In 2008, a team of Italian chemists found a way to remove the polymer coatings without damaging the underlying fresco. They treated the murals with a microemulsion, a special type of mixture where molecules called surfactants encapsulate and carry nanometer-sized droplets of organic solvents through an aqueous solution to the surface of the fresco. Once the droplets reach the surface, they can dissolve and carry away the unwanted polymer coating.

1.6 | 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



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

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 1.24** 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 1.24 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 ± 0.0001 g. The ± notation (read "plus or minus") expresses the magnitude of the uncertainty of your measurement. In much scientific work, we drop the ± notation with the understanding that *there is always some uncertainty in the last digit reported for any measured quantity.*

Figure 1.25 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 °C 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 °C or 26 °C, but we can't say that it is *exactly* 27 °C.

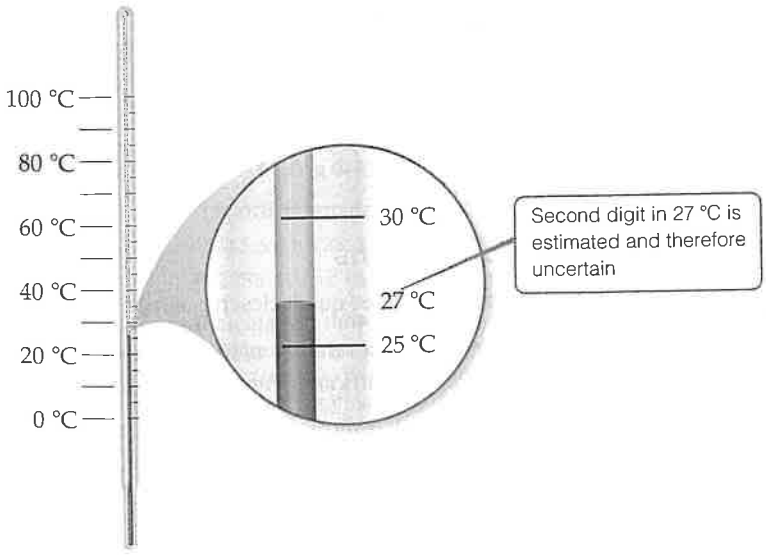
All digits of a measured quantity, including the uncertain one, are called **significant figures**. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures. The greater the number of significant figures, the greater the precision implied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

Give It Some Thought

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?

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:

- Zeros *between* nonzero digits are always significant—1005 kg (four significant figures); 7.03 cm (three significant figures).
- 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).
- 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).



▲ Figure 1.25 Uncertainty and significant figures in a measurement.

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 1.24 Precision and accuracy.

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 5,546,574 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. Moreover, 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 5,500,000 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**
An object is determined to have a mass of 0.01080 g. How many significant figures are there in this measurement?
(**a**) 2 (**b**) 3 (**c**) 4 (**d**) 5 (**e**) 6

▶ **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

Apply the following rule 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 the number of the significant figures in the result

20.42	← two decimal places
1.322	← three decimal places
83.1	← one decimal places
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.

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.

When *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

The mass of a glass beaker is known to be 25.1 g. Approximately 5 mL of water are added, and the mass of the beaker and water is measured on an analytical balance to be 30.625 g. How many significant figures are there in the mass of the water?

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.

*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 1.9

Determining the Number of Significant Figures in a Calculated Quantity

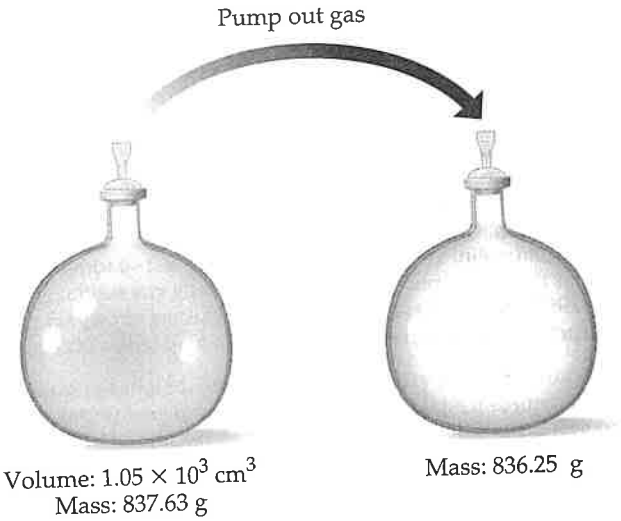
A vessel containing a gas at 25 °C is weighed, emptied, and then reweighed as depicted in **Figure 1.26**. From the data provided, calculate the density of the gas at 25 °C.

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:

$$\begin{array}{r} 837.63\ \text{g} \\ -836.25\ \text{g} \\ \hline 1.38\ \text{g} \end{array}$$

In subtracting numbers, we determine the number of significant figures in our result by counting decimal places in each quantity.



▲ **Figure 1.26 Uncertainty and significant figures in a measurement.**

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.

▶ **Practice Exercise 1**

You are asked to determine the mass of a piece of copper using its reported density, 8.96 g/mL, and a 150-mL graduated cylinder. First, you add 105 mL of water to the graduated cylinder; then you place the piece of copper in the cylinder and record a volume of 137 mL. What is the mass of the copper reported with the correct number of significant figures? **(a)** 287 g **(b)** 3.5×10^{-3} g/mL **(c)** 286.72 g/mL **(d)** 3.48×10^{-3} g/mL **(e)** 2.9×10^2 g/mL

▶ **Practice Exercise 2**

If the mass of the container in the sample exercise (Figure 1.26) 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.

1.7 | 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.

Conversion Factors

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:

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

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

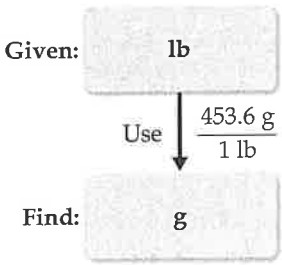
If a woman has a mass of 115 lb, what is her mass in grams? (Use the relationships between units given on the back inside cover of the text.)

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:

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

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 on the top right coulmn.



▶ **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 the back inside cover for the 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 FOR SUCCESS 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

Consider the following conversion factors: **(i)** 1000 m = 1 km, **(ii)** 1 mL = 1 cm^3, **(iii)** 1 lb = 453.59 g. Which of these conversion factors has the potential to change the number of significant figures in a calculation?

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

Given: m/s → (Use $\frac{1\text{ km}}{10^3\text{ m}}$) → km/s → (Use $\frac{1\text{ mi}}{1.6093\text{ km}}$) → mi/s → (Use $\frac{60\text{ s}}{1\text{ min}}$) → mi/min → (Use $\frac{60\text{ min}}{1\text{ hr}}$) → Find: mi/hr

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

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

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

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 $0.3 \times 60 = 20$ mi/min. Multiplying again by 60 gives $20 \times 60 = 1200$ 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.

Given: in.^3 → (Use $\left(\frac{2.54\text{ cm}}{1\text{ in.}}\right)^3$) → cm^3 → (Use $\frac{19.3\text{ g}}{1\text{ cm}^3}$) → Find: g

Sample Exercise 1.12

Converting Volume Units

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

SOLUTION

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

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

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

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

$$= 1.36 \times 10^{21}\ L$$



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

▶ **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 $510 \times 10^6\ km^2$, 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 FOR SUCCESS 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 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 master the content of the chapters more fully 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.

STRATEGIES FOR SUCCESS 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 also describe how to take advantage of the text's website, 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.

- Exercises for each section of a chapter are grouped in pairs, with answers given at the back of the book for the odd-numbered (red) ones.
- 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 in 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 those of biological interest.
- WebElements (http://www.webelements.com/) is a good website 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.

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:

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

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



A composite deck.

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 kg × 2.2 lb/kg = 8.3 lb, is thus reasonable.

▶ **Practice Exercise 1**

Composite decking is a manufactured substitute for wood compounded from post-consumer plastic and wood. It is frequently used in outdoor decks. The density of a particular composite decking is reported as 60.0 lb/ft^3. What is the density in kg/L?

(a) 138 kg/L **(b)** 0.961 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.

Chapter Summary and Key Terms

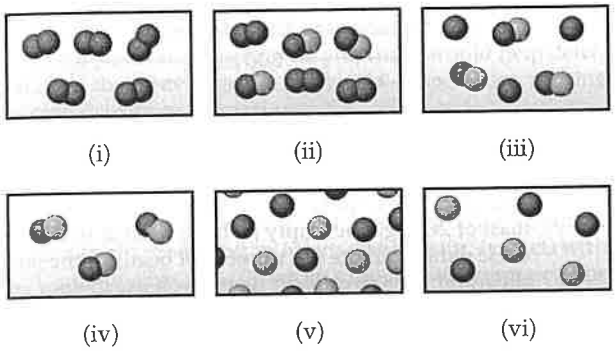
THE STUDY OF CHEMISTRY (SECTION 1.1) **Chemistry** is the study of the composition, structure, properties, and changes of **matter**. The composition of matter relates to the kinds of **elements** it contains. The structure of matter relates to the ways the **atoms** of these elements are arranged. A **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.

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.

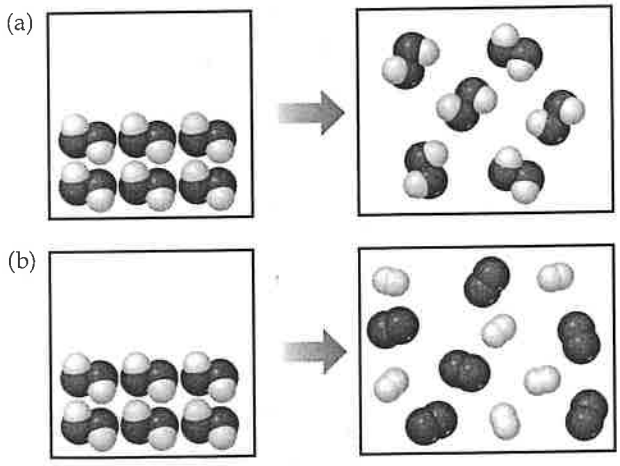
Exercises

Visualizing Concepts

1.1 Which of the following figures represents (**a**) a pure element, (**b**) a mixture of two elements, (**c**) a pure compound, (**d**) a mixture of an element and a compound? (More than one picture might fit each description.) [Section 1.2]



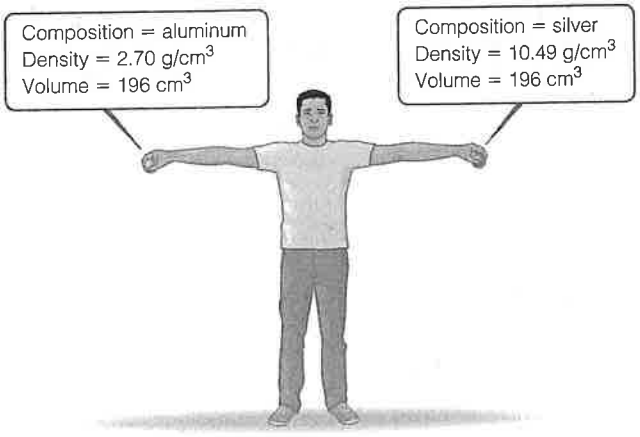
1.2 Which of the following diagrams represents a chemical change? [Section 1.3]



1.3 Musical instruments like trumpets and trombones are made from an alloy called brass. Brass is composed of copper and zinc atoms and appears homogeneous under an optical microscope. The approximate composition of most brass objects is a 2:1 ratio of copper to zinc atoms, but the exact ratio varies somewhat from one piece of brass to another. (**a**) Would you classify brass as an element, a compound, a homogeneous mixture, or a heterogeneous mixture? (**b**) Would it be correct to say that brass is a solution? [Section 1.2]



1.4 Consider the two spheres shown here, one made of silver and the other of aluminum. (**a**) What is the mass of each sphere in kg? (**b**) The force of gravity acting on an object is $F = mg$, where m is the mass of an object and g is the acceleration of gravity (9.8 m/s^2). How much work do you do on each sphere it you raise it from the floor to a height of 2.2 m? (**c**) Does the act of lifting the sphere off the ground increase the potential energy of the aluminum sphere by a larger, smaller, or same amount as the silver sphere? (**d**) If you release the spheres simultaneously, they will have the same velocity when they hit the ground. Will they have the same kinetic energy? If not, which sphere will have more kinetic energy? [Section 1.4]



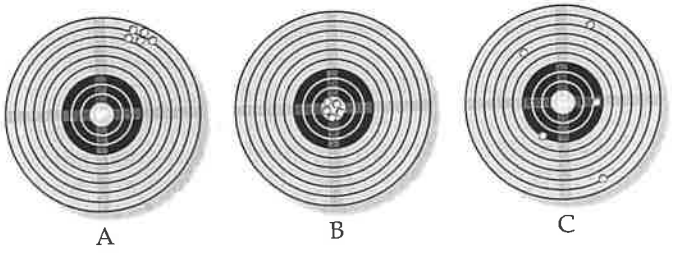
1.5 Is the separation method used in brewing a cup of coffee best described as distillation, filtration, or chromatography? [Section 1.3]



1.6 Identify each of the following as measurements of length, area, volume, mass, density, time, or temperature: (**a**) 25 ps, (**b**) 374.2 mg, (**c**) 77 K, (**d**) 100,000 km^2, (**e**) 1.06 μm, (**f**) 16 nm^2, (**g**) −78 °C, (**h**) 2.56 g/cm^3, (**i**) 28 cm^3. [Section 1.5]

1.7 (**a**) Three spheres of equal size are composed of aluminum (density = 2.70 g/cm^3), silver (density = 10.49 g/cm^3), and nickel (density = 8.90 g/cm^3). List the spheres from lightest to heaviest. (**b**) Three cubes of equal mass are composed of gold (density = 19.32 g/cm^3), platinum (density = 21.45 g/cm^3), and lead (density = 11.35 g/cm^3). List the cubes from smallest to largest. [Section 1.5]

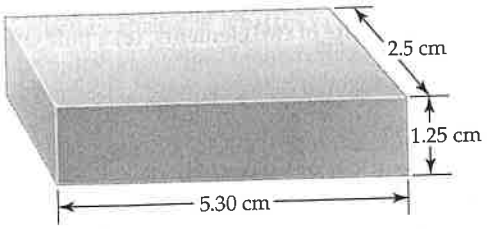
1.8 The three targets from a rifle range shown below were produced by: (**A**) the instructor firing a newly acquired target rifle; (**B**) the instructor firing his personal target rifle; and (**C**) a student who has fired his target rifle only a few times. (**a**) Comment on the accuracy and precision for each of these three sets of results. (**b**) For the A and C results in the future to look like those in B, what needs to happen? [Section 1.6]



1.9 (**a**) What is the length of the pencil in the following figure if the ruler reads in centimeters? How many significant figures are there in this measurement? (**b**) An automobile speedometer with circular scales reading both miles per hour and kilometers per hour is shown. What speed is indicated, in both units? How many significant figures are in the measurements? [Section 1.6]



1.10 (**a**) How many significant figures should be reported for the volume of the metal bar shown here? (**b**) If the mass of the bar is 104.72 g, how many significant figures should be reported when its density is determined using the calculated volume? [Section 1.6]



1.11 Consider the jar of jelly beans in the photo. To get an estimate of the number of beans in the jar you weigh six beans and obtain masses of 3.15, 3.12, 2.98, 3.14, 3.02, and 3.09 g. Then you weigh the jar with all the beans in it, and obtain a mass of 2082 g. The empty jar has a mass of 653 g. Based on these data, estimate the number of beans in the jar. Justify the number of significant figures you use in your estimate. [Section 1.6]



1.12 The photo below shows a picture of an agate stone. Jack, who picked up the stone on the Lake Superior shoreline and polished it, insists that agate is a chemical compound. Ellen argues that it cannot be a compound. Discuss the relative merits of their positions. [Section 1.2]



Classification and Properties of Matter (Sections 1.2 and 1.3)

1.13 Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (**a**) rice pudding, (**b**) seawater, (**c**) magnesium, (**d**) crushed ice.

1.14 Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (**a**) air, (**b**) tomato juice, (**c**) iodine crystals, (**d**) sand.

1.15 Give the chemical symbol or name for the following elements, as appropriate: (**a**) sulfur, (**b**) gold, (**c**) potassium, (**d**) chlorine, (**e**) copper, (**f**) U, (**g**) Ni, (**h**) Na, (**i**) Al, (**j**) Si.

1.16 Give the chemical symbol or name for each of the following elements, as appropriate: (**a**) carbon, (**b**) nitrogen, (**c**) titanium, (**d**) zinc, (**e**) iron, (**f**) P, (**g**) Ca, (**h**) He, (**i**) Pb, (**j**) Ag.

1.17 A solid white substance A is heated strongly in the absence of air. It decomposes to form a new white substance B and a gas C. The gas has exactly the same properties as the product obtained when carbon is burned in an excess of oxygen. Based on these observations, can we determine whether solids A and B and gas C are elements or compounds?

1.18 You are hiking in the mountains and find a shiny gold nugget. It might be the element gold, or it might be "fool's gold," which is a nickname for iron pyrite, FeS_2. Which of the following physical properties do you think would help determine if the shiny nugget is really gold—appearance, melting point, density, or physical state?

1.19 In the process of attempting to characterize a substance, a chemist makes the following observations: The substance is a silvery white, lustrous metal. It melts at 649 °C and boils at 1105 °C. Its density at 20 °C is 1.738 g/cm^3. The substance burns in air, producing an intense white light. It reacts with chlorine to give a brittle white solid. The substance can be pounded into thin sheets or drawn into wires. It is a good conductor of electricity. Which of these characteristics are physical properties, and which are chemical properties?

1.20 (**a**) Read the following description of the element zinc and indicate which are physical properties and which are chemical properties.



Zinc melts at 420 °C. When zinc granules are added to dilute sulfuric acid, hydrogen is given off and the metal dissolves. Zinc has a hardness on the Mohs scale of 2.5 and a density of 7.13 g/cm^3 at 25 °C. It reacts slowly with oxygen gas at elevated temperatures to form zinc oxide, ZnO.

(**b**) Which properties of zinc can you describe from the photo? Are these physical or chemical properties?

1.21 Label each of the following as either a physical process or a chemical process: (**a**) rusting of a metal can, (**b**) boiling a cup of water, (**c**) pulverizing an aspirin, (**d**) digesting a candy bar, (**e**) exploding of nitroglyerin.

1.22 A match is lit and held under a cold piece of metal. The following observations are made: (**a**) The match burns. (**b**) The metal gets warmer. (**c**) Water condenses on the metal. (**d**) Soot (carbon) is deposited on the metal. Which of these occurrences are due to physical changes, and which are due to chemical changes?

1.23 Which separation method is better suited for separating a solution of sugar and water into pure substances, filtration or distillation?

1.24 Two beakers contain clear, colorless liquids. When the contents of the beakers are mixed a white solid is formed. (**a**) Is this an example of a chemical or a physical change? (**b**) What would be the most convenient way to separate the newly formed white solid from the liquid mixture—filtration, distillation, or chromatography.

The Nature of Energy (Section 1.4)

1.25 (**a**) Calculate the kinetic energy, in joules of a 1200-kg automobile moving at 18 m/s. (**b**) Convert this energy to calories. (**c**) When the automobile brakes to a stop is the "lost" kinetic energy converted mostly to heat or to some form of potential energy?

1.26 (**a**) A baseball weighs 5.13 oz. What is the kinetic energy, in joules, of this baseball when it is thrown by a major league pitcher at 95.0 mi/h? (**b**) By what factor will the kinetic energy change if the speed of the baseball is decreased to 55.0 mi/h? (**c**) What happens to the kinetic energy when the baseball is caught by the catcher? Is it converted mostly to heat or to some form of potential energy?

1.27 Two positively charged particles are first brought close together and then released. Once released, the repulsion between particles causes them to move away from each other. (**a**) This is an example of potential energy being converted into what form of energy? (**b**) Does the potential energy of the two particles prior to release increase or decrease as the distance between them is increased.

1.28 For each of the following processes, does the potential energy of the object(s) increase or decrease? (**a**) The distance between two oppositely charged particles is increased. (**b**) Water is pumped from ground level to the reservoir of a water tower 30 m above the ground. (**c**) The bond in a chlorine molecule, Cl_2, is broken to form two chlorine atoms.

1.29 What is the kinetic energy and velocity of the aluminum sphere in Problem 1.4 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

1.30 What is the kinetic energy and velocity of the silver sphere in Problem 1.4 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

Units of Measurement (Section 1.5)

1.31 What exponential notation do the following abbreviations represent? (**a**) d, (**b**) c, (**c**) f, (**d**) μ, (**e**) M, (**f**) k, (**g**) n, (**h**) m, (**i**) p.

1.32 Use appropriate metric prefixes to write the following measurements without use of exponents: (**a**) 2.3×10^{-10} L, (**b**) 4.7×10^{-6} g, (**c**) 1.85×10^{-12} m, (**d**) 16.7×10^{6} s, (**e**) 15.7×10^{3} g, (**f**) 1.34×10^{-3} m, (**g**) 1.84×10^{2} cm.

1.33 Make the following conversions: (**a**) 72 °F to °C, (**b**) 216.7 °C to °F, (**c**) 233 °C to K, (**d**) 315 K to °F, (**e**) 2500 °F to K, (**f**) 0 K to °F.

1.34 (**a**) The temperature on a warm summer day is 87 °F. What is the temperature in °C? (**b**) Many scientific data are reported at 25 °C. What is this temperature in kelvins and in degrees Fahrenheit? (**c**) Suppose that a recipe calls for an oven temperature of 400 °F. Convert this temperature to degrees Celsius and to kelvins. (**d**) Liquid nitrogen boils at 77 K. Convert this temperature to degrees Fahrenheit and to degrees Celsius.

1.35 (**a**) A sample of tetrachloroethylene, a liquid used in dry cleaning that is being phased out because of its potential to cause cancer, has a mass of 40.55 g and a volume of 25.0 mL at 25 °C. What is its density at this temperature? Will tetrachloroethylene float on water? (Materials that are less dense than water will float.) (**b**) Carbon dioxide (CO_2) is a gas at room temperature and pressure. However, carbon dioxide can be put under pressure to become a "supercritical fluid" that is a much safer dry-cleaning agent than tetrachloroethylene. At a certain pressure, the density of supercritical CO_2 is 0.469 g/cm^3. What is the mass of a 25.0-mL sample of supercritical CO_2 at this pressure?

1.36 (**a**) A cube of osmium metal 1.500 cm on a side has a mass of 76.31 g at 25 °C. What is its density in g/cm^3 at this temperature? (**b**) The density of titanium metal is 4.51 g/cm^3 at 25 °C. What mass of titanium displaces 125.0 mL of water at 25 °C?

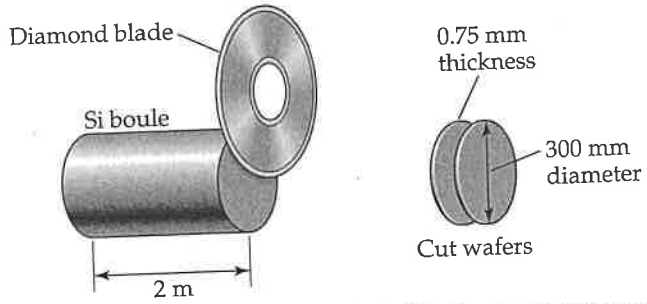
(c) The density of benzene at 15 °C is 0.8787 g/mL. Calculate the mass of 0.1500 L of benzene at this temperature.

1.37 (a) To identify a liquid substance, a student determined its density. Using a graduated cylinder, she measured out a 45-mL sample of the substance. She then measured the mass of the sample, finding that it weighed 38.5 g. She knew that the substance had to be either isopropyl alcohol (density 0.785 g/mL) or toluene (density 0.866 g/mL). What are the calculated density and the probable identity of the substance? (b) An experiment requires 45.0 g of ethylene glycol, a liquid whose density is 1.114 g/mL. Rather than weigh the sample on a balance, a chemist chooses to dispense the liquid using a graduated cylinder. What volume of the liquid should he use? (c) Is a graduated cylinder such as that shown in Figure 1.21 likely to afford the accuracy of measurement needed? (d) A cubic piece of metal measures 5.00 cm on each edge. If the metal is nickel, whose density is 8.90 g/cm^3, what is the mass of the cube?

1.38 (a) After the label fell off a bottle containing a clear liquid believed to be benzene, a chemist measured the density of the liquid to verify its identity. A 25.0-mL portion of the liquid had a mass of 21.95 g. A chemistry handbook lists the density of benzene at 15 °C as 0.8787 g/mL. Is the calculated density in agreement with the tabulated value? (b) An experiment requires 15.0 g of cyclohexane, whose density at 25 °C is 0.7781 g/mL. What volume of cyclohexane should be used? (c) A spherical ball of lead has a diameter of 5.0 cm. What is the mass of the sphere if lead has a density of 11.34 g/cm^3? (The volume of a sphere is $(4/3)\pi r^3$, where r is the radius.)

1.39 In the year 2013, an estimated amount of 36 billion metric tons (1 metric ton = 1000 kg) of carbon dioxide (CO_2) was emitted worldwide due to fossil fuel combustion and cement production. Express this mass of CO_2 in grams without exponential notation, using an appropriate metric prefix.

1.40 Silicon for computer chips is grown in large cylinders called "boules" that are 300 mm in diameter and 2 m in length, as shown. The density of silicon is 2.33 g/cm^3. Silicon wafers for making integrated circuits are sliced from a 2.0-m boule and are typically 0.75 mm thick and 300 mm in diameter. (a) How many wafers can be cut from a single boule? (b) What is the mass of a silicon wafer? (The volume of a cylinder is given by $\pi r^2 h$, where r is the radius and h is its height.)



1.41 Use of the British thermal unit (Btu) is common in some types of engineering work. A Btu is the amount of heat required to raise the temperature of 1 lb of water by 1°F. Calculate the number of joules in a Btu.

1.42 A watt is a measure of power (the rate of energy change) equal to 1 J/s. (a) Calculate the number of joules in a kilowatt-hour. (b) An adult person radiates heat to the surroundings at about the same rate as a 100-watt electric incandescent light bulb. What is the total amount of energy in kcal radiated to the surroundings by an adult over a 24 h period?

Uncertainty in Measurement (Section 1.6)

1.43 Indicate which of the following are exact numbers: (a) the mass of a 3- by 5-in. index card, (b) the number of ounces in a pound, (c) the volume of a cup of Seattle's Best coffee, (d) the number of inches in a mile, (e) the number of microseconds in a week, (f) the number of pages in this book.

1.44 Indicate which of the following are exact numbers: (a) the mass of a 32-oz can of coffee, (b) the number of students in your chemistry class, (c) the temperature of the surface of the Sun, (d) the mass of a postage stamp, (e) the number of milliliters in a cubic meter of water, (f) the average height of NBA basketball players.

1.45 What is the number of significant figures in each of the following measured quantities? (a) 601 kg, (b) 0.054 s, (c) 6.3050 cm, (d) 0.0105 L, (e) 7.0500×10^{-3} m^3, (f) 400 g.

1.46 Indicate the number of significant figures in each of the following measured quantities: (a) 3.774 km, (b) 205 m^2, (c) 1.700 cm, (d) 350.00 K, (e) 307.080 g, (f) 1.3×10^3 m/s.

1.47 Round each of the following numbers to four significant figures and express the result in standard exponential notation: (a) 102.53070, (b) 656.980, (c) 0.008543210, (d) 0.000257870, (e) −0.0357202.

1.48 (a) The diameter of Earth at the equator is 7926.381 mi. Round this number to three significant figures and express it in standard exponential notation. (b) The circumference of Earth through the poles is 40,008 km. Round this number to four significant figures and express it in standard exponential notation.

1.49 Carry out the following operations and express the answers with the appropriate number of significant figures.

(a) 14.3505 + 2.65 (b) 952.7 − 140.7389

(c) $(3.29 \times 10^4)(0.2501)$ (d) 0.0588/0.677

1.50 Carry out the following operations and express the answer with the appropriate number of significant figures.

(a) 320.5 − (6104.5/2.3)

(b) $[(285.3 \times 10^5) - (1.200 \times 10^3)] \times 2.8954$

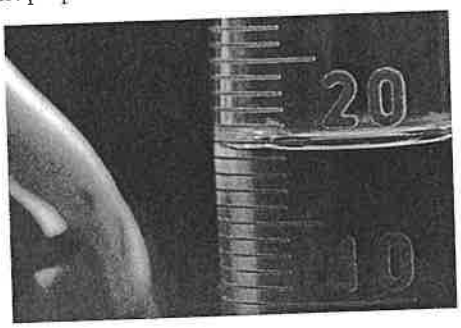
(c) $(0.0045 \times 20{,}000.0) + (2813 \times 12)$

(d) $863 \times [1255 - (3.45 \times 108)]$

1.51 You weigh an object on a balance and read the mass in grams according to the picture. How many significant figures are in this measurement?



1.52 You have a graduated cylinder that contains a liquid (see photograph). Write the volume of the liquid, in milliliters, using the proper number of significant figures.



Dimensional Analysis (Section 1.7)

1.53 Using your knowledge of metric units, English units, and the information on the back inside cover, write down the conversion factors needed to convert (**a**) mm to nm, (**b**) mg to kg, (**c**) km to ft, (**d**) $in.^3$ to cm^3.

1.54 Using your knowledge of metric units, English units, and the information on the back inside cover, write down the conversion factors needed to convert (**a**) μm to mm, (**b**) ms to ns, (**c**) mi to km, (**d**) ft^3 to L.

1.55 (**a**) A bumblebee flies with a ground speed of 15.2 m/s. Calculate its speed in km/hr. (**b**) The lung capacity of the blue whale is 5.0×10^3 L. Convert this volume into gallons. (**c**) The Statue of Liberty is 151 ft tall. Calculate its height in meters. (**d**) Bamboo can grow up to 60.0 cm/day. Convert this growth rate into inches per hour.

1.56 (**a**) The speed of light in a vacuum is 2.998×10^8 m/s. Calculate its speed in miles per hour. (**b**) The Sears Tower in Chicago is 1454 ft tall. Calculate its height in meters. (**c**) The Vehicle Assembly Building at the Kennedy Space Center in Florida has a volume of 3,666,500 m^3. Convert this volume to liters and express the result in standard exponential notation. (**d**) An individual suffering from a high cholesterol level in her blood has 242 mg of cholesterol per 100 mL of blood. If the total blood volume of the individual is 5.2 L, how many grams of total blood cholesterol does the individual's body contain?

1.57 Perform the following conversions: (**a**) 5.00 days to s, (**b**) 0.0550 mi to m, (**c**) \$1.89/gal to dollars per liter, (**d**) 0.510 in./ms to km/hr, (**e**) 22.50 gal/min to L/s, (**f**) 0.02500 ft^3 to cm^3.

1.58 Carry out the following conversions: (**a**) 0.105 in. to mm, (**b**) 0.650 qt to mL, (**c**) 8.75 μm/s to km/hr, (**d**) 1.955 m^3 to yd^3 (**e**) \$3.99/lb to dollars per kg, (**f**) 8.75 lb/ft^3 to g/mL.

1.59 (**a**) How many liters of wine can be held in a wine barrel whose capacity is 31 gal? (**b**) The recommended adult dose of Elixophyllin®, a drug used to treat asthma, is 6 mg/kg of body mass. Calculate the dose in milligrams for a 185-lb person. (**c**) If an automobile is able to travel 400 km on 47.3 L of gasoline, what is the gas mileage in miles per gallon? (**d**) When the coffee is brewed according to directions, a pound of coffee beans yields 50 cups of coffee (4 cups = 1 qt). How many kg of coffee are required to produce 200 cups of coffee?

1.60 (**a**) If an electric car is capable of going 225 km on a single charge, how many charges will it need to travel from Seattle, Washington, to San Diego, California, a distance of 1257 mi, assuming that the trip begins with a full charge? (**b**) If a migrating loon flies at an average speed of 14 m/s, what is its average speed in mi/hr? (**c**) What is the engine piston displacement in liters of an engine whose displacement is listed as 450 $in.^3$? (**d**) In March 1989, the *Exxon Valdez* ran aground and spilled 240,000 barrels of crude petroleum off the coast of Alaska. One barrel of petroleum is equal to 42 gal. How many liters of petroleum were spilled?

1.61 The density of air at ordinary atmospheric pressure and 25 °C is 1.19 g/L. What is the mass, in kilograms, of the air in a room that measures 14.5 ft × 16.5 ft × 8.0 ft?

1.62 The concentration of carbon monoxide in an urban apartment is 48 $\mu g/m^3$. What mass of carbon monoxide in grams is present in a room measuring 10.6 ft × 14.8 ft × 20.5 ft?

1.63 Gold can be hammered into extremely thin sheets called gold leaf. An architect wants to cover a 100 ft × 82 ft ceiling with gold leaf that is five-millionths of an inch thick. The density of gold is 19.32 g/cm^3, and gold costs \$1654 per troy ounce (1 troy ounce = 31.1034768 g). How much will it cost the architect to buy the necessary gold?

1.64 A copper refinery produces a copper ingot weighing 150 lb. If the copper is drawn into wire whose diameter is 7.50 mm, how many feet of copper can be obtained from the ingot? The density of copper is 8.94 g/cm^3. (Assume that the wire is a cylinder whose volume $V = \pi r^2 h$, where r is its radius and h is its height or length.)

Additional Exercises

1.65 Classify each of the following as a pure substance, a solution, or a heterogeneous mixture: (**a**) a gold ingot, (**b**) a cup of coffee, (**c**) a wood plank.

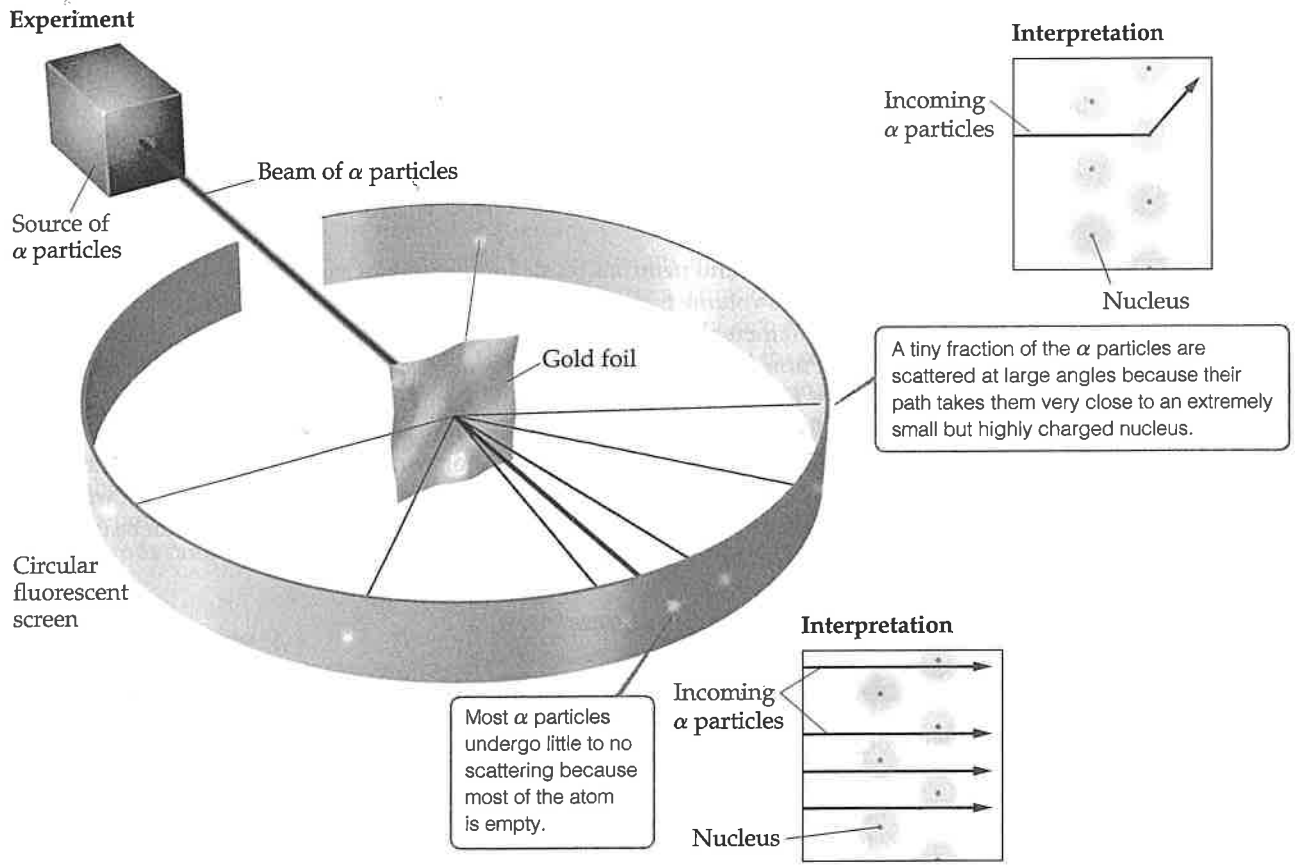
1.66 (**a**) Which is more likely to eventually be shown to be incorrect: an hypothesis or a theory? (**b**) A(n) _____ reliably predicts the behavior of matter, while a(n) _____ provides an explanation for that behavior.

1.67 A sample of ascorbic acid (vitamin C) is synthesized in the laboratory. It contains 1.50 g of carbon and 2.00 g of oxygen. Another sample of ascorbic acid isolated from citrus fruits contains 6.35 g of carbon. According to the law of constant composition, how many grams of oxygen does it contain?

1.68 Ethyl chloride is sold as a liquid (see photo) under pressure for use as a local skin anesthetic. Ethyl chloride boils at 12 °C at atmospheric pressure. When the liquid is sprayed onto the skin, it boils off, cooling and numbing the skin as it vaporizes. (**a**) What changes of state are involved in this use of ethyl chloride? (**b**) What is the boiling point of ethyl chloride in degrees Fahrenheit? (**c**) The bottle shown contains 103.5 mL of ethyl chloride. The density of ethyl chloride at 25 °C is 0.765 g/cm^3. What is the mass of ethyl chloride in the bottle?



Go Figure What is the charge on the particles that form the beam? Will they be attracted to or repelled from the positively charged gold nuclei?



▲ **Figure 2.9 Rutherford's α-scattering experiment.** When α particles pass through a gold foil, most pass through undeflected but some are scattered, a few at very large angles. According to the plum-pudding model of the atom, the particles should experience only very minor deflections. The nuclear model of the atom explains why a few α particles are deflected at large angles. Although the nuclear atom has been depicted here as a yellow sphere, it is important to realize that most of the space around the nucleus contains only the low-mass electrons.

repulsion between the highly positive charge of the gold nucleus and the positive charge of the α particle was strong enough to deflect the particle, as shown in Figure 2.9.

Subsequent experiments led to the discovery of positive particles (**protons**) and neutral particles (**neutrons**) in the nucleus. Protons were discovered in 1919 by Rutherford and neutrons in 1932 by British scientist James Chadwick (1891–1972). Thus, the atom is composed of electrons, protons, and neutrons.

Give It Some Thought

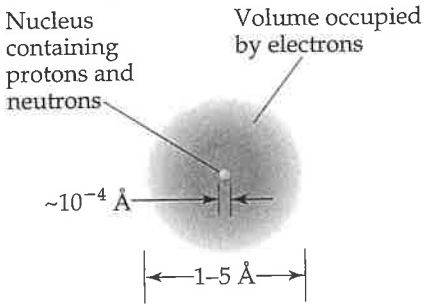
Would you expect the number of α particles scattered at large angles in Rutherford's experiment to increase, decrease, or stay the same as the thickness of the gold foil increases?

2.3 The Modern View of Atomic Structure

Since Rutherford's time, as physicists have learned more and more about atomic nuclei, the list of particles that make up nuclei has grown and continues to increase. You may have come across the names of some subatomic particles, such as quarks, leptons, and

Go Figure

What is the approximate diameter of the nucleus in units of pm?



▲ **Figure 2.10 The structure of the atom.** A cloud of rapidly moving electrons occupies most of the volume of the atom. The nucleus occupies a tiny region at the center of the atom and is composed of the protons and neutrons. The nucleus contains virtually all the mass of the atom.

bosons.* As chemists, however, we can take a simple view of the atom because only three subatomic particles—the proton, neutron, and electron—have a bearing on chemical behavior.

As noted earlier, the charge of an electron is -1.602×10^{-19} C. The charge of a proton is opposite in sign but equal in magnitude to that of an electron: $+1.602 \times 10^{-19}$ C. The quantity 1.602×10^{-19} C is called the **electronic charge**. For convenience, the charges of atomic and subatomic particles are usually expressed as multiples of this charge rather than as coulombs. Thus, the charge of an electron is 1– and that of a proton is 1+. Neutrons are electrically neutral (which is how they received their name). *Every atom has an equal number of electrons and protons, so atoms have no net electrical charge.*

Protons and neutrons reside in the tiny nucleus of the atom. The vast majority of an atom's volume is the space in which the electrons reside (**Figure 2.10**). Most atoms have diameters between 1×10^{-10} m (100 pm) and 5×10^{-10} m (500 pm). A convenient non-SI unit of length used for atomic dimensions is the **angstrom** (Å), where $1 \text{ Å} = 1 \times 10^{-10}$ m = 100 pm. Thus, atoms have diameters of approximately 1 – 5 Å. The diameter of a chlorine atom, for example, is 200 pm, or 2.0 Å.

Electrons are attracted to the protons in the nucleus by the electrostatic force that exists between particles of opposite electrical charge. In later chapters, we will see that the strength of the attractive forces between electrons and nuclei can be used to explain many of the differences among different elements.

Give It Some Thought

(a) If an atom has 15 protons, how many electrons does it have?
(b) Where do the protons reside in an atom?

Atoms have extremely small masses. The mass of the heaviest known atom, for example, is approximately 4×10^{-22} g. Because it would be cumbersome to express such small masses in grams, we use the **atomic mass unit** (amu),** where $1 \text{ amu} = 1.66054 \times 10^{-24}$ g. A proton has a mass of 1.0073 amu, a neutron 1.0087 amu, and an electron 5.486×10^{-4} amu (**Table 2.1**). Because it takes 1836 electrons to equal the mass of one proton, and 1839 electrons to equal the mass of a single neutron, the nucleus accounts for nearly the entire mass of an atom.

TABLE 2.1 Comparison of the Proton, Neutron, and Electron

Particle	Charge	Mass (amu)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1–)	5.486×10^{-4}

The diameter of an atomic nucleus is approximately 10^{-4} Å, only a small fraction of the diameter of the atom as a whole. You can appreciate the relative sizes of the atom and its nucleus by imagining that if the hydrogen atom were as large as a football stadium, the nucleus would be the size of a small marble. Because the tiny nucleus carries most of the mass of the atom in such a small volume, it has an incredibly high density—on the order of 10^{13}–10^{14} g/cm^3. A matchbox full of material of such density would weigh over 2.5 billion tons!

Figure 2.10 incorporates the features we have just discussed. Electrons play the major role in chemical reactions. The significance of representing the region containing

*The electron is an elementary particle that cannot be divided into smaller particles, whereas protons and neutrons are made up of smaller particles called quarks.

**The SI abbreviation for the atomic mass unit is u. We will use the more common abbreviation amu.

electrons as an indistinct cloud will become clear in later chapters when we consider the energies and spatial arrangements of the electrons. For now, however, we have all the information we need to discuss many topics that form the basis of everyday uses of chemistry.

Atomic Numbers, Mass Numbers, and Isotopes

What makes an atom of one element different from an atom of another element? The atoms of each element have a *characteristic number of protons*. The number of protons in an atom of any particular element is called that element's **atomic number**. Because an atom has no net electrical charge, the number of electrons it contains must equal the number of protons. All atoms of carbon, for example, have six protons and six electrons, whereas all atoms of oxygen have eight protons and eight electrons. Thus, carbon has atomic number 6, and oxygen has atomic number 8. The atomic number of each element is listed with the name and symbol of the element on the front inside cover of the text.

Atoms of a given element can differ in the number of neutrons they contain and, consequently, in mass. For example, while most atoms of carbon have six neutrons, some have more and some have less. The symbol $^{12}_{6}C$ (read "carbon twelve," carbon-12) represents the carbon atom containing six protons and six neutrons, whereas carbon atoms that contain six protons and eight neutrons have mass number 14, are represented as $^{14}_{6}C$ and are referred to as carbon-14.

Sample Exercise 2.1

Atomic Size

The diameter of a U.S. dime is 17.9 mm, and the diameter of a silver atom is 2.88 Å. How many silver atoms could be arranged side by side across the diameter of a dime?

SOLUTION

The unknown is the number of silver (Ag) atoms. Using the relationship 1 Ag atom = 2.88 Å as a conversion factor relating number of atoms and distance, we start with the diameter of the dime, first converting this distance into angstroms and then using the diameter of the Ag atom to convert distance to number of Ag atoms:

$$\text{Ag atoms} = (17.9\ \cancel{\text{mm}})\left(\frac{10^{-3}\ \cancel{\text{m}}}{1\ \cancel{\text{mm}}}\right)\left(\frac{1\ \text{Å}}{10^{-10}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{Ag atom}}{2.88\ \text{Å}}\right)$$

$$= 6.22 \times 10^{7}\ \text{Ag atoms}$$

That is, 62.2 million silver atoms could sit side by side across a dime!

▶ **Practice Exercise 1**

Which of the following factors determines the size of an atom? (**a**) the volume of the nucleus (**b**) the volume of space occupied by the electrons of the atom (**c**) the volume of a single electron, multiplied by the number of electrons in the atom (**d**) the total nuclear charge (**e**) the total mass of the electrons surrounding the nucleus

▶ **Practice Exercise 2**

The diameter of a carbon atom is 1.54 Å. (**a**) Express this diameter in picometers. (**b**) How many carbon atoms could be aligned side by side across the width of a pencil line that is 0.20 mm wide?

A CLOSER LOOK Basic Forces

Four basic forces are known in nature: (1) gravitational, (2) electromagnetic, (3) strong nuclear, and (4) weak nuclear. *Gravitational forces* are attractive forces that act between all objects in proportion to their masses. Gravitational forces between atoms or between subatomic particles are so small that they are of no chemical significance.

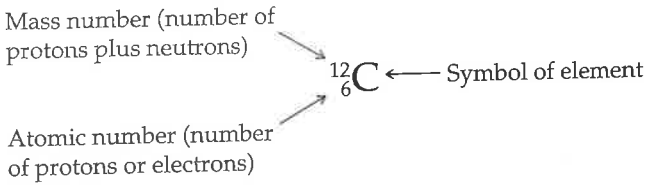
Electromagnetic forces are attractive or repulsive forces that act between either electrically charged or magnetic objects. The magnitude of the electric force between two charged particles is given by *Coulomb's law*: $F = kQ_1Q_2/d^2$, where Q_1 and Q_2 are the magnitudes of the charges on the two particles, d is the distance between their centers, and k is a constant determined by the units for Q and d. ⟲ (Section 1.4) A negative value for the force indicates attraction, whereas a positive value indicates repulsion. Electric forces are of primary importance in determining the chemical properties of elements.

All nuclei except those of hydrogen atoms contain two or more protons. Because like charges repel, electrical repulsion would cause the protons to fly apart if the *strong nuclear force* did not keep them together. As the name implies, this force can be quite strong but only when particles are extremely close together, as are the protons and neutrons in a nucleus. At this distance, the attractive strong nuclear force is stronger than the positive–positive repulsive electric force and holds the nucleus together.

The *weak nuclear force* is weaker than the electric force and the strong nuclear force but stronger than the gravitational force. We are aware of its existence only because it shows itself in certain types of radioactivity.

Related Exercise: 2.114

The atomic number is indicated by the subscript; the superscript, called the **mass number**, is the number of protons plus neutrons in the atom:



Because all atoms of a given element have the same atomic number, the subscript is redundant and is often omitted. Thus, the symbol for carbon-12 can be represented simply as ^{12}C.

Atoms with identical atomic numbers but different mass numbers (that is, the same number of protons but different numbers of neutrons) are called **isotopes** of one another. Several isotopes of carbon are listed in **Table 2.2**. We will generally use the notation with superscripts only when referring to a particular isotope of an element. It is important to keep in mind that the isotopes of any given element are all alike chemically. A carbon dioxide molecule that contains a ^{13}C atom behaves for all practical purposes identically to one that contains a ^{12}C atom.

TABLE 2.2 Some Isotopes of Carbon[a]

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
^{11}C	6	6	5
^{12}C	6	6	6
^{13}C	6	6	7
^{14}C	6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Sample Exercise 2.2

Determining the Number of Subatomic Particles in Atoms

How many protons, neutrons, and electrons are in an atom of (**a**) ^{197}Au, (**b**) strontium-90?

SOLUTION

(**a**) The superscript 197 is the mass number (protons + neutrons). According to the list of elements given on the front inside cover, gold has atomic number 79. Consequently, an atom of ^{197}Au has 79 protons, 79 electrons, and 197 − 79 = 118 neutrons. (**b**) The atomic number of strontium is 38. Thus, all atoms of this element have 38 protons and 38 electrons. The strontium-90 isotope has 90 − 38 = 52 neutrons.

▶ **Practice Exercise 1**

Which of these atoms has the largest number of neutrons? (**a**) ^{148}Eu (**b**) ^{157}Dy (**c**) ^{149}Nd (**d**) ^{162}Ho (**e**) ^{159}Gd

▶ **Practice Exercise 2**

How many protons, neutrons, and electrons are in an atom of (**a**) ^{138}Ba, (**b**) phosphorus-31?

Sample Exercise 2.3

Writing Symbols for Atoms

Magnesium has three isotopes with mass numbers 24, 25, and 26. (**a**) Write the complete chemical symbol (superscript and subscript) for each. (**b**) How many neutrons are in an atom of each isotope?

SOLUTION

(**a**) Magnesium has atomic number 12, so all atoms of magnesium contain 12 protons and 12 electrons. The three isotopes are therefore represented by $^{24}_{12}Mg$, $^{25}_{12}Mg$, and $^{26}_{12}Mg$. (**b**) The number of neutrons in each isotope is the mass number minus the number of protons. The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

▶ **Practice Exercise 1**

Which of the following is an incorrect representation for a neutral atom? (**a**) $^{6}_{3}Li$ (**b**) $^{13}_{6}C$ (**c**) $^{63}_{30}Cu$ (**d**) $^{30}_{15}P$ (**e**) $^{108}_{47}Ag$

▶ **Practice Exercise 2**

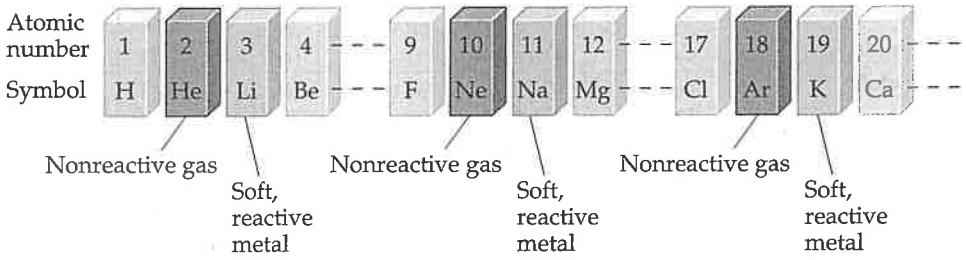
Give the complete chemical symbol for the atom that contains 82 protons, 82 electrons, and 126 neutrons.

2.5 | The Periodic Table

As the list of known elements expanded during the early 1800s, attempts were made to find patterns in chemical behavior. These efforts culminated in the development of the periodic table in 1869. We will have much to say about the periodic table in later chapters, but it is so important and useful that you should become acquainted with it now. You will quickly learn that *the periodic table is the most significant tool that chemists use for organizing and remembering chemical facts.*

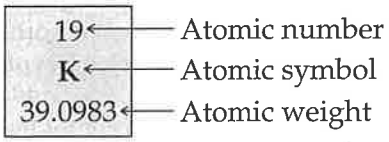
Many elements show strong similarities to one another. The elements lithium (Li), sodium (Na), and potassium (K) are all soft, very reactive metals, for example. The elements helium (He), neon (Ne), and argon (Ar) are all nonreactive gases. If the elements are arranged in order of increasing atomic number, their chemical and physical properties show a repeating, or *periodic*, pattern. For example, each of the soft, reactive metals—lithium, sodium, and potassium—comes immediately after one of the nonreactive gases—helium, neon, and argon, respectively—as shown in **Figure 2.13.**

Go Figure If F is a reactive nonmetal, which other element or elements shown here are likely to be reactive nonmetals?



▲ Figure 2.13 Arranging elements by atomic number reveals a periodic pattern of properties. This pattern is the basis of the periodic table.

The arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns, is known as the **periodic table** (**Figure 2.14**). The table shows the atomic number and atomic symbol for each element, and the atomic weight is often given as well, as in this typical entry for potassium:

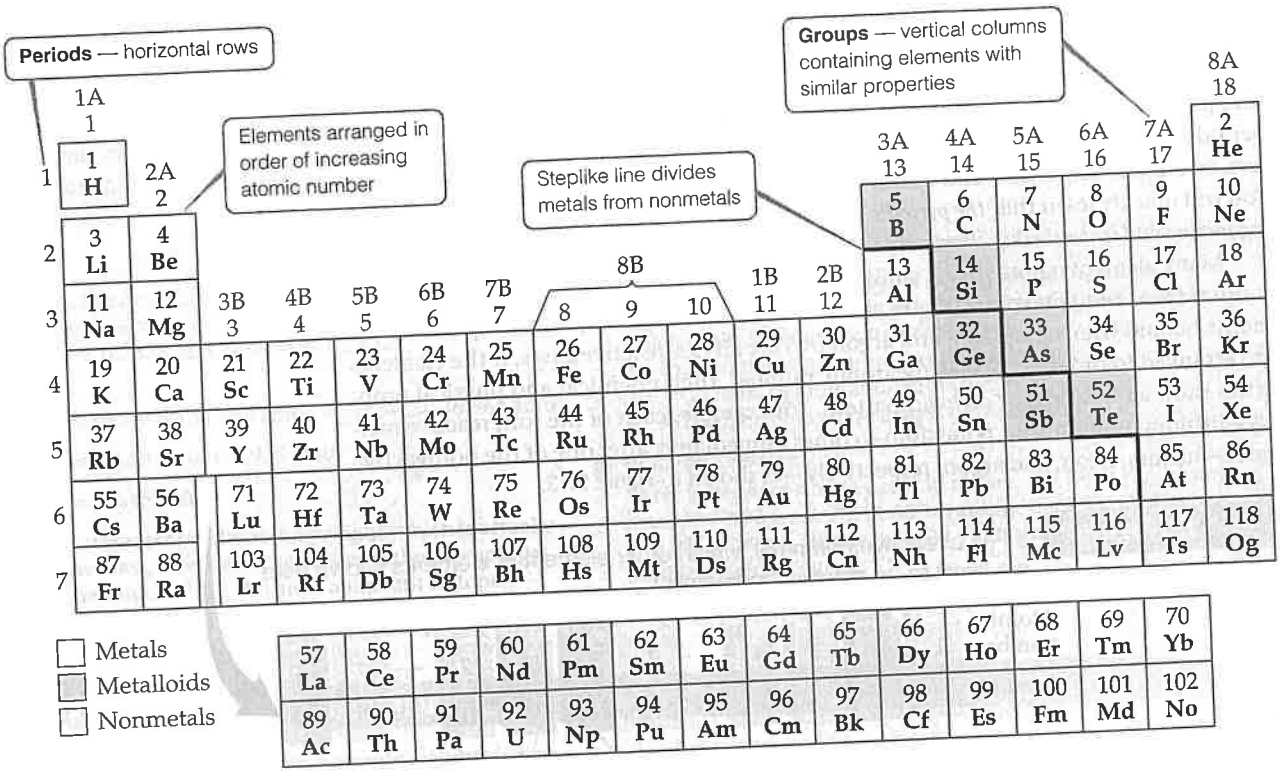


You might notice slight variations in periodic tables from one book to another or between those in the lecture hall and in the text. These are simply matters of style, or they might concern the particular information included. There are no fundamental differences.

The horizontal rows of the periodic table are called **periods**. The first period consists of only two elements, hydrogen (H) and helium (He). The second and third periods consist of eight elements each. The fourth and fifth periods contain 18 elements. The sixth and seventh periods have 32 elements each, but in order to fit on a page 14 of the elements from each period (atomic numbers 57–70 and 89–102) appear at the bottom of the table.

The vertical columns are **groups**. The way in which the groups are labeled is somewhat arbitrary. Three labeling schemes are in common use, two of which are shown in Figure 2.14.

- The top set of labels, which have A and B designations, is widely used in North America. Roman numerals, rather than Arabic ones, are often employed in this scheme. Group 7A, for example, is often labeled VIIA.
- Europeans use a similar convention that assigns the A and B labels differently.
- In an effort to eliminate confusion, the International Union of Pure and Applied Chemistry (IUPAC) has proposed a convention that numbers the groups from 1 through 18 with no A or B designations, as shown in Figure 2.14.



▲ **Figure 2.14** Periodic table of elements.

We will use the traditional North American convention with Arabic numerals and the letters A and B.

Elements in a group often exhibit similarities in physical and chemical properties. For example, the "coinage metals"—copper (Cu), silver (Ag), and gold (Au)—belong to group 1B. These elements are less reactive than most metals, which is why they have been traditionally used throughout the world to make coins. Many other groups in the periodic table also have names, listed in Table 2.3.

We will learn in Chapters 6 and 7 that elements in a group have similar properties because they have the same arrangement of electrons at the periphery of their atoms. However, we need not wait until then to make good use of the periodic table; after all, the chemists who developed the table knew nothing about electrons! We can use the table, as they intended, to correlate behaviors of elements and to help us remember many facts.

The color code of Figure 2.14 shows that, except for hydrogen, all the elements on the left and in the middle of the table are **metallic elements**, or **metals**. All the metallic elements share characteristic properties, such as luster and high electrical and heat conductivity, and all of them except mercury (Hg) are solid at room temperature.*

TABLE 2.3 Names of Some Groups in the Periodic Table

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens	O, S, Se, Te, Po
7A	Halogens	F, Cl, Br, I, At
8A	Noble gases	He, Ne, Ar, Kr, Xe, Rn

*All metals become liquids if heated sufficiently. Hg simply has the lowest melting point of any metallic element. Although sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and gallium (Ga) are solids at room temperature, they all melt at temperatures below 100 °C.

A CLOSER LOOK What Are Coins Made Of?

Copper, silver, and gold were traditionally employed to make coins, but modern coins are typically made from other metals (**Figure 2.15**). To be useful for coinage, a metal, or combination of metals (called an *alloy*), must be corrosion resistant. It must also be hard enough to withstand rough usage and yet be of a consistency that permits machines to accurately stamp the coins. Some metals that might otherwise make fine coins—for example, manganese (Mn)—are ruled out because they make the coins too hard to stamp. A third criterion is that the value of the metal in the coin should not be as great as the face value of the coin.

This last criterion has led to several changes in the composition of U.S. coins over the past century. The U.S. Mint ceased producing gold coins in 1933 during the Great Depression. In 1964, a silver crisis caused the replacement of silver in quarters and dimes. Finally, in 1982, the composition of the penny was changed from nearly pure copper to copper-plated zinc. If pennies were made today from pure copper, the metal would be worth more than a penny, thus inviting smelters to melt down the coins for the value of the metal.

One of the traditional alloys for making coins is a mixture of copper and nickel. Today only the U.S. nickel is made from this alloy, called cupronickel, which consists of 75% copper and 25% nickel. The modern U.S. dollar coin, often referred to as the silver dollar, doesn't contain any silver. It consists of copper (88.5%), zinc (6.0%), manganese (3.5%), and nickel (2.0%).

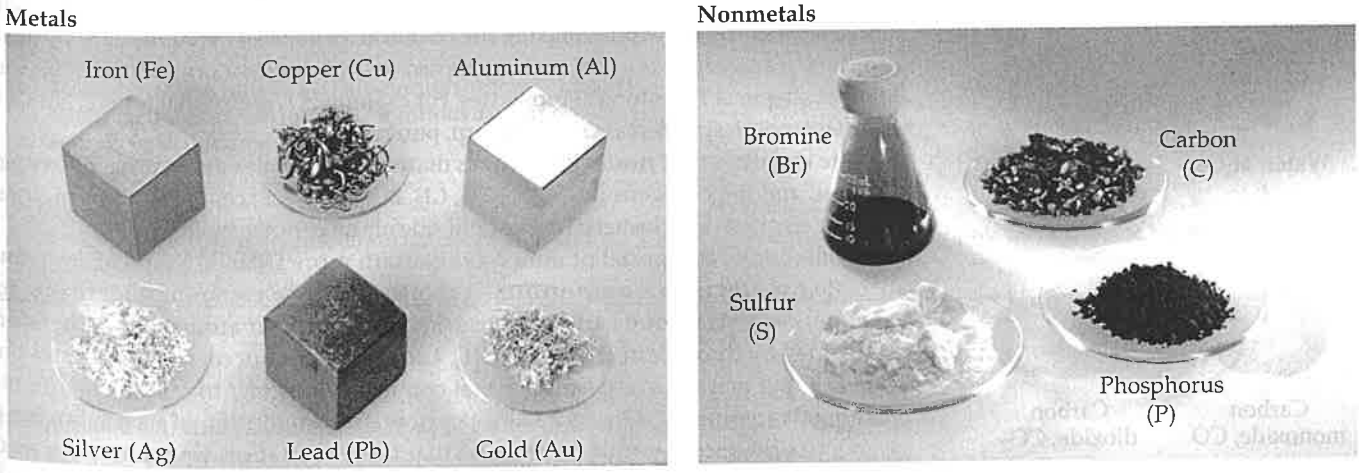
Related Exercises: 2.103, 2.104



▲ Figure 2.15 The US quarter dollar is made from an alloy that is 91.67% Cu and 8.33% Ni.

The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a stepped line that runs from boron (B) to astatine (At). (Note that hydrogen, although on the left side of the table, is a nonmetal.) At room temperature and pressure, some of the nonmetals are gaseous, some are solid, and one is liquid. Nonmetals generally differ from metals in appearance (**Figure 2.16**) and in other physical properties. Many of the elements that lie along the line that separates metals from nonmetals have properties that fall between those of metals and nonmetals. These elements are often referred to as **metalloids**.

Go Figure Name two ways in which the metals shown here differ in general appearance from the nonmetals.



▲ Figure 2.16 Examples of metals and nonmetals.

Give It Some Thought

Chlorine is a halogen (Table 2.3). Locate this element in the periodic table.

(a) What is its symbol?
(b) In which period and in which group is the element located?
(c) What is its atomic number?
(d) Is it a metal or nonmetal?

Sample Exercise 2.5

Using the Periodic Table

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

SOLUTION

Elements in the same group of the periodic table are most likely to exhibit similar properties. We therefore expect Ca and Mg to be most alike because they are in the same group (2A, the alkaline earth metals).

▶ **Practice Exercise 1**

A biochemist who is studying the properties of certain sulfur (S)–containing compounds in the body wonders whether trace amounts of another nonmetallic element might have similar behavior. To which element should she turn her attention? (**a**) F (**b**) As (**c**) Se (**d**) Cr (**e**) P

▶ **Practice Exercise 2**

Locate Na (sodium) and Br (bromine) in the periodic table. Give the atomic number of each and classify each as metal, metalloid, or nonmetal.

2.6 | Molecules and Molecular Compounds

Even though the atom is the smallest representative sample of an element, only the noble-gas elements are normally found in nature as isolated atoms. Most matter is composed of molecules or ions. We examine molecules here and ions in Section 2.7.

Molecules and Chemical Formulas

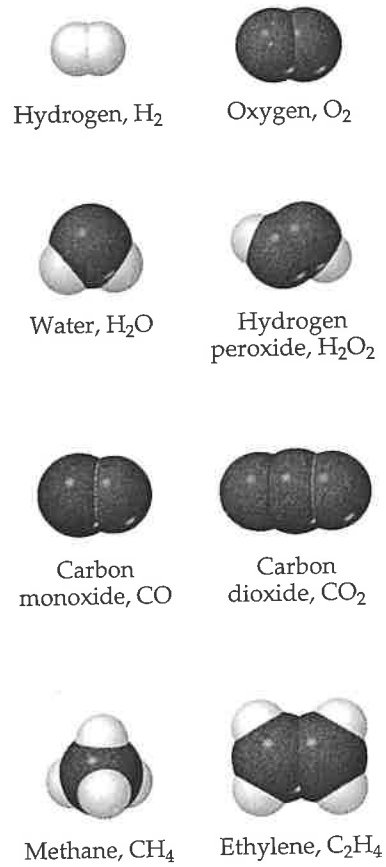
Several elements are found in nature in molecular form—two or more of the same type of atom bound together. For example, most of the oxygen in air consists of molecules that contain two oxygen atoms. As we saw in Section 1.2, we represent this molecular oxygen by the **chemical formula** O_2 (read "oh two"). The subscript tells us that two oxygen atoms are present in each molecule. A molecule made up of two atoms is called a **diatomic molecule**.

Oxygen also exists in another molecular form known as *ozone*. Molecules of ozone consist of three oxygen atoms, making the chemical formula O_3. Even though "normal" oxygen (O_2) and ozone (O_3) are both composed only of oxygen atoms, they exhibit very different chemical and physical properties. For example, O_2 is essential for life, but O_3 is toxic; O_2 is odorless, whereas O_3 has a sharp, pungent smell.

The elements that normally occur as diatomic molecules are hydrogen, oxygen, nitrogen, and the halogens (H_2, O_2, N_2, F_2, Cl_2, Br_2, and I_2). Except for hydrogen, these diatomic elements are clustered on the right side of the periodic table.

Compounds composed of molecules contain more than one type of atom and are called **molecular compounds**. A molecule of the compound methane, for example, consists of one carbon atom and four hydrogen atoms and is therefore represented by the chemical formula CH_4. Lack of a subscript on the C indicates one atom of C per methane molecule. Several common molecules of both elements and compounds are shown in **Figure 2.17.** Notice how the composition of each substance is given by its chemical formula. Notice also that these substances are composed only of nonmetallic elements.

Most of the molecular substances we will encounter contain only nonmetals.



▲ **Figure 2.17 Molecular models.** Notice how the chemical formulas of these simple molecules correspond to their compositions.

Molecular and Empirical Formulas

Chemical formulas that indicate the actual numbers of atoms in a molecule are called **molecular formulas.** (The formulas in Figure 2.17 are molecular formulas.) Chemical formulas that give only the relative number of atoms of each type in a molecule are called **empirical formulas**. The subscripts in an empirical formula are always the smallest possible whole-number ratios. The molecular formula for hydrogen peroxide is H_2O_2, for example, whereas its empirical formula is HO. The molecular formula for ethylene is

C_2H_4, and its empirical formula is CH_2. For many substances, the molecular formula and the empirical formula are identical, as in the case of water, H_2O.

Give It Some Thought

Consider the following four formulas: SO_2, B_2H_6, CO, $C_4H_2O_2$. Which of these formulas could be (**a**) only an empirical formula, (**b**) only a molecular formula, (**c**) either a molecular or an empirical formula?

Whenever we know the molecular formula of a compound, we can determine its empirical formula. The converse is not true, however. If we know the empirical formula of a substance, we cannot determine its molecular formula unless we have more information. So why do chemists bother with empirical formulas? As we will see in Chapter 3, certain common methods of analyzing substances lead to the empirical formula only. For example, if you decomposed hydrogen peroxide H_2O_2 into its elements and weighed them, you could determine that there were equal numbers of hydrogen and oxygen atoms, but you would not know if the molecular formula was HO, H_2O_2, H_3O_3, or the like. Once the empirical formula is known, additional experiments can give the information needed to convert the empirical formula to the molecular one. In addition, there are many substances that do not exist as isolated molecules, one example being ionic compounds that are discussed later in this chapter. For these substances, we must rely on empirical formulas.

Sample Exercise 2.6

Relating Empirical and Molecular Formulas

Write the empirical formulas for (**a**) glucose, a substance also known as either blood sugar or dextrose—molecular formula $C_6H_{12}O_6$; (**b**) nitrous oxide, a substance used as an anesthetic and commonly called laughing gas—molecular formula N_2O.

SOLUTION

(**a**) The subscripts of an empirical formula are the smallest whole-number ratios. The smallest ratios are obtained by dividing each subscript by the largest common factor, in this case 6. The resultant empirical formula for glucose is CH_2O.

(**b**) Because the subscripts in N_2O are already the lowest integral numbers, the empirical formula for nitrous oxide is the same as its molecular formula, N_2O.

▶ **Practice Exercise 1**

Tetracarbon dioxide is an unstable oxide of carbon with the following molecular structure:



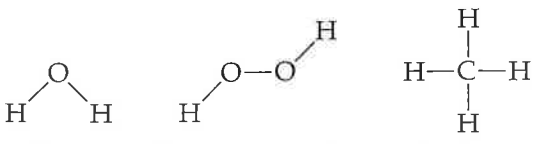
What are the molecular and empirical formulas of this substance? (**a**) C_2O_2, CO_2 (**b**) C_4O, CO (**c**) CO_2, CO_2 (**d**) C_4O_2, C_2O (**e**) C_2O, CO_2

▶ **Practice Exercise 2**

Give the empirical formula for *decaborane*, whose molecular formula is $B_{10}H_{14}$.

Picturing Molecules

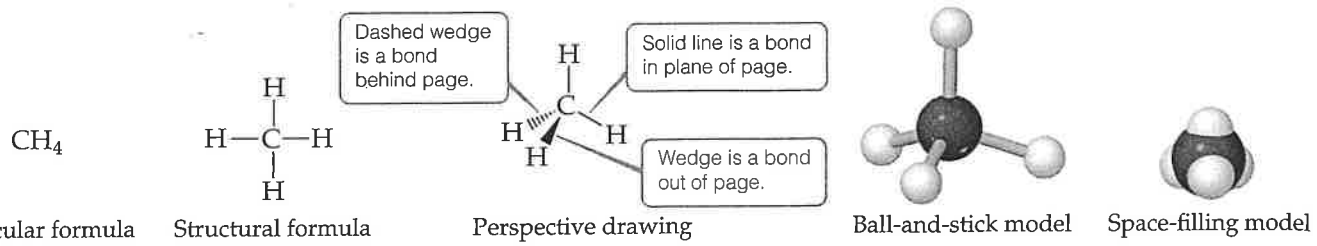
The molecular formula of a substance does not show how its atoms are joined together. A **structural formula** is needed to convey that information, as in the following examples:



Water Hydrogen peroxide Methane

The atoms are represented by their chemical symbols, and lines are used to represent the bonds that hold the atoms together.

Go Figure Which model, the ball-and-stick or the space-filling, more effectively shows the angles between bonds around a central atom?



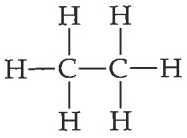
▲ **Figure 2.18 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models.

A structural formula does not typically depict the actual geometry of the molecule, that is, the actual angles at which atoms are joined; for that, more sophisticated representations are needed (**Figure 2.18**).

- **Perspective drawings** use wedges and dashed lines to depict bonds that are not in the plane of the paper. This gives a crude sense of the three-dimensional shape of a molecule.
- **Ball-and-stick models** show atoms as spheres and bonds as sticks. This type of model has the advantage of accurately representing the angles at which the atoms are attached to one another in a molecule (Figure 2.18). Sometimes the chemical symbols of the elements are superimposed on the balls, but often the atoms are identified simply by color.
- **Space-filling models** depict what a molecule would look like if the atoms were scaled up in size (Figure 2.18). These models show the relative sizes of the atoms, but the angles between atoms, which help define their molecular geometry, are often more difficult to see than in ball-and-stick models. Because space-filling models give a good representation of the true size of a molecule, they are useful for picturing how two molecules might fit together or pack in the solid state. As with ball-and-stick models, the identities of the atoms are typically indicated by color.

Give It Some Thought

The structural formula for ethane is

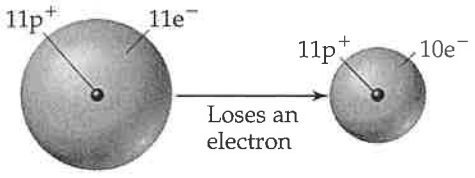


(a) What is the molecular formula for ethane?
(b) What is its empirical formula?
(c) Can we infer from this drawing that the H—C—H bond angles are 90 degrees?

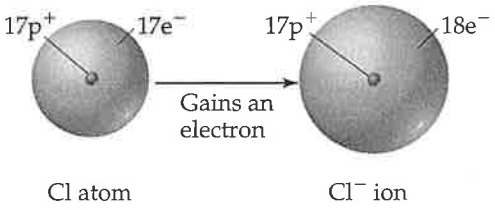
2.7 | Ions and Ionic Compounds

If electrons are removed from or added to an atom, a charged particle called an **ion** is formed. An ion with a positive charge is a **cation** (pronounced CAT-ion); a negatively charged ion is an **anion** (AN-ion).

To see how ions form, consider the sodium atom, which has 11 protons and 11 electrons. This atom easily loses one electron. The resulting cation has 11 protons and 10 electrons, which means it has a net charge of 1+.



The net charge on an ion is represented by a superscript. The superscripts +, 2+, and 3+, for instance, mean a net charge resulting from the *loss* of one, two, and three electrons, respectively. The superscripts −, 2−, and 3− represent net charges resulting from the *gain* of one, two, and three electrons, respectively. Chlorine, with 17 protons and 17 electrons, for example, can gain an electron in chemical reactions, producing the Cl^- ion:



In general, metal atoms tend to lose electrons to form cations and nonmetal atoms tend to gain electrons to form anions. Thus, ionic compounds tend to be composed of both metal cations and nonmetal anions, as in NaCl.

Sample Exercise 2.7

Writing Chemical Symbols for Ions

Give the chemical symbol, including superscript indicating mass number, for (**a**) the ion with 22 protons, 26 neutrons, and 19 electrons; and (**b**) the ion of sulfur that has 16 neutrons and 18 electrons.

SOLUTION

(**a**) The number of protons is the atomic number of the element. A periodic table or list of elements tells us that the element with atomic number 22 is titanium (Ti). The mass number (protons plus neutrons) of this isotope of titanium is 22 + 26 = 48. Because the ion has three more protons than electrons, it has a net charge of 3+ and is designated $^{48}Ti^{3+}$.

(**b**) The periodic table tells us that sulfur (S) has an atomic number of 16. Thus, each atom or ion of sulfur contains 16 protons. We are told that the ion also has 16 neutrons, meaning the mass number is 16 + 16 = 32. Because the ion has 16 protons and 18 electrons, its net charge is 2− and the ion symbol is $^{32}S^{2-}$.

In general, we will focus on the net charges of ions and ignore their mass numbers unless the circumstances dictate that we specify a certain isotope.

▶ **Practice Exercise 1**

In which of the following species is the difference between the number of protons and the number of electrons largest? (**a**) Ti^{2+} (**b**) P^{3-} (**c**) Mn (**d**) Se^{2-} (**e**) Ce^{4+}

▶ **Practice Exercise 2**

How many protons, neutrons, and electrons does the $^{79}Se^{2-}$ ion possess?

In addition to simple ions such as Na^+ and Cl^-, there are **polyatomic ions**, such as NH_4^+ (ammonium ion) and SO_4^{2-} (sulfate ion), which consist of atoms joined as in a molecule, but carrying a net positive or negative charge. Polyatomic ions will be discussed in Section 2.8.

It is important to realize that the chemical properties of ions are very different from the chemical properties of the atoms from which the ions are derived. The addition or removal of one or more electrons produces a charged species with behavior very different from that of its associated atom or group of atoms. For example, sodium metal reacts violently with water, but an ionic compound containing sodium ions, such as NaCl, does not.

Predicting Ionic Charges

As noted in Table 2.3, the elements of group 8A are called the noble-gas elements. The noble gases are chemically nonreactive elements that form very few compounds. Many atoms gain or lose electrons to end up with the same number of electrons as the noble

gas closest to them in the periodic table. We might deduce that atoms tend to acquire the electron arrangements of the noble gases because these electron arrangements are very stable. Nearby elements can obtain these same stable arrangements by losing or gaining electrons. For example, the loss of one electron from an atom of sodium leaves it with the same number of electrons as in a neon atom (10). Similarly, when chlorine gains an electron, it ends up with 18, the same number of electrons as in argon. This simple observation will be helpful for now to account for the formation of ions. A deeper explanation awaits us in Chapter 8, where we discuss chemical bonding.

Sample Exercise 2.8

Predicting Ionic Charge

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

SOLUTION

We will assume that barium and oxygen form ions that have the same number of electrons as the nearest noble-gas atom. From the periodic table, we see that barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the Ba^{2+} cation.

Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the O^{2-} anion.

▶ Practice Exercise 1

Although it is helpful to know that many ions have the electron arrangement of a noble gas, many elements, especially among the metals, form ions that do not have a noble-gas electron arrangement. Use the periodic table, Figure 2.14, to determine which of the following ions has a noble-gas electron arrangement, and which do not. For those that do, indicate the noble-gas arrangement they match: **(a)** Ti^{4+}, **(b)** Mn^{2+}, **(c)** Pb^{2+}, **(d)** Te^{2-}, **(e)** Zn^{2+}.

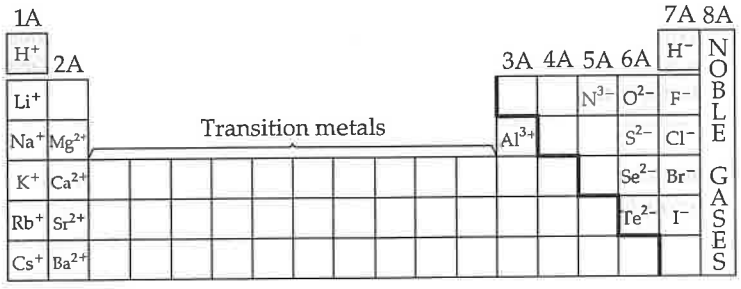
▶ Practice Exercise 2

Predict the charge expected for the most stable ion of **(a)** aluminum and **(b)** fluorine.

The periodic table is very useful for remembering ionic charges, especially those of elements on the left and right sides of the table. As **Figure 2.19** shows, the charges of these ions relate in a simple way to their positions in the table: The group 1A elements (alkali metals) form 1+ ions, the group 2A elements (alkaline earths) form 2+ ions, the group 7A elements (halogens) form 1− ions, and the group 6A elements form 2− ions. As noted in Practice Exercise 1 of Sample Exercise 2.8, many of the other groups do not lend themselves to such simple rules.

Go Figure

The most common ions for silver, zinc, and scandium are Ag^+, Zn^{2+}, and Sc^{3+}. Locate the boxes in which you would place these ions in this table. Which of these ions has the same number of electrons as a noble-gas element?



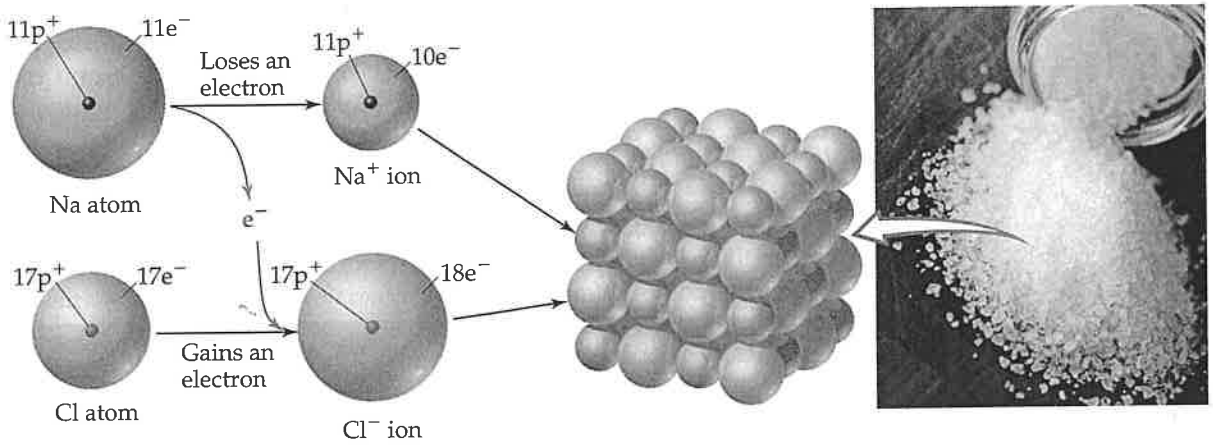
▲ **Figure 2.19 Predictable charges of some common ions.** Notice that the red stepped line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.

Ionic Compounds

A great deal of chemical activity involves the transfer of electrons from one substance to another. **Figure 2.20** shows that when elemental sodium is allowed to react with elemental chlorine, an electron transfers from a sodium atom to a chlorine atom, forming a Na^+ ion and a Cl^- ion. Because objects of opposite charges attract, the Na^+ and the Cl^- ions bind together to form the compound sodium chloride (NaCl). Sodium chloride,

which we know better as common table salt, is an example of an **ionic compound**, a compound made up of cations and anions.

Go Figure Is it possible to pick out a unique cluster of atoms that can be thought of as a sodium chloride molecule?



▲ **Figure 2.20 Formation of an ionic compound.** The transfer of an electron from a sodium atom to a chlorine atom leads to the formation of a Na^+ ion and a Cl^- ion. These ions are arranged in a lattice in solid sodium chloride, NaCl.

We can often tell whether a compound is ionic (consisting of ions) or molecular (consisting of molecules) from its composition. As a general rule, cations are metal ions and anions are nonmetal ions. Consequently,

Ionic compounds are generally combinations of metals and nonmetals, as in NaCl.

Sample Exercise 2.9

Identifying Ionic and Molecular Compounds

Which of these compounds would you expect to be ionic: N_2O, Na_2O, $CaCl_2$, SF_4?

SOLUTION

We predict that Na_2O and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal. We predict (correctly) that N_2O and SF_4 are molecular compounds because they are composed entirely of nonmetals.

▶ **Practice Exercise 1**

Which of these compounds are molecular: CBr_4, FeS, P_4O_6, PbF_2?

▶ **Practice Exercise 2**

Give a reason why each of the following statements is a safe prediction:

(**a**) Every compound of Rb with a nonmetal is ionic in character.
(**b**) Every compound of nitrogen with a halogen element is a molecular compound.
(**c**) The compound $MgKr_2$ does not exist.
(**d**) Na and K are very similar in the compounds they form with nonmetals.
(**e**) If contained in an ionic compound, calcium (Ca) will be in the form of the doubly charged ion, Ca^{2+}.

The ions in ionic compounds are arranged in three-dimensional structures, as Figure 2.20 shows for NaCl. Because there is no discrete "molecule" of NaCl, we are able to write only an empirical formula for this substance. This is true for most other ionic compounds.

We can write the empirical formula for an ionic compound if we know the charges of the ions. Because chemical compounds are always electrically neutral, the ions in an ionic compound always occur in such a ratio that the total positive charge equals the total negative charge. Thus, there is one Na^+ to one Cl^- in NaCl, one Ba^{2+} to two Cl^- in $BaCl_2$, and so forth.

As you consider these and other examples, you will see that if the charges on the cation and anion are equal, the subscript on each ion is 1. If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion. For

example, the ionic compound formed from Mg (which forms Mg^{2+} ions) and N (which forms N^{3-} ions) is Mg_3N_2:



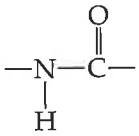
There is one caveat to using this approach. Remember that the empirical formula should be the smallest possible whole-number ratio of the two elements. So the empirical formula for the ionic compound formed between Ti^{4+} and O^{2-} is TiO_2 rather than Ti_2O_4.

Give It Some Thought

What is the empirical formula of the compound formed between W^{6+} and O^{2-}?

CHEMISTRY AND LIFE Elements Required by Living Organisms

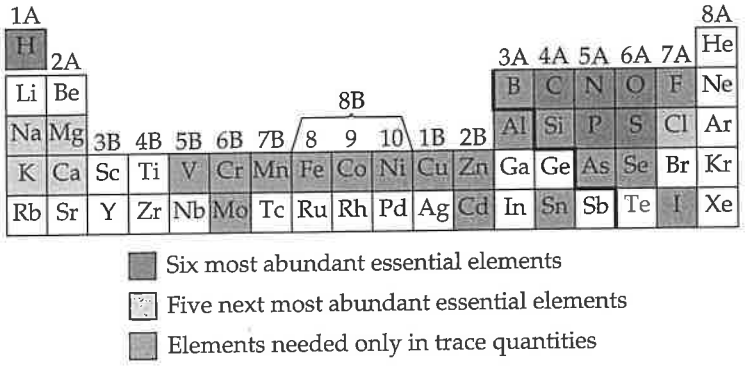
The elements essential to life are highlighted in color in **Figure 2.21.** More than 97% of the mass of most organisms is made up of just six of these elements—oxygen, carbon, hydrogen, nitrogen, phosphorus, and sulfur. Water is the most common compound in living organisms, accounting for at least 70% of the mass of most cells. In the solid components of cells, carbon is the most prevalent element by mass. Carbon atoms are found in a vast variety of organic molecules, bonded either to other carbon atoms or to atoms of other elements. Nearly all proteins, for example, contain the carbon-based group



which occurs repeatedly in the molecules.

In addition, 23 other elements have been found in various living organisms. Five are ions required by all organisms: Ca^{2+}, Cl^-, Mg^{2+}, K^+, and Na^+. Calcium ions, for example, are necessary for the formation of bone and transmission of nervous system signals. Many other elements are needed in only very small quantities and consequently are called *trace* elements. For example, trace quantities of copper are required in the diet of humans to aid in the synthesis of hemoglobin.

Related Exercise: 2.102



▲ **Figure 2.21 Elements essential to life.**

Sample Exercise 2.10

Using Ionic Charge to Write Empirical Formulas for Ionic Compounds

Write the empirical formula of the compound formed by (**a**) Al^{3+} and Cl^- ions, (**b**) Al^{3+} and O^{2-} ions, (**c**) Mg^{2+} and NO_3^- ions.

SOLUTION

(**a**) Three Cl^- ions are required to balance the charge of one Al^{3+} ion, making the empirical formula $AlCl_3$.

(**b**) Two Al^{3+} ions are required to balance the charge of three O^{2-} ions. A 2:3 ratio is needed to balance the total positive charge of 6+ and the total negative charge of 6–. The empirical formula is Al_2O_3.

(**c**) Two NO_3^- ions are needed to balance the charge of one Mg^{2+}, yielding $Mg(NO_3)_2$. Note that the formula for the polyatomic ion, NO_3^-, must be enclosed in parentheses so that it is clear that the subscript 2 applies to all the atoms of that ion.

▶ Practice Exercise 1

Which of the following nonmetals will form an ionic compound with Sc^{3+} that has a 1:1 ratio of cations to anions?
(**a**) Ne (**b**) F (**c**) O (**d**) N

▶ Practice Exercise 2

Write the empirical formula for the compound formed by (**a**) Na^+ and PO_4^{3-}, (**b**) Zn^{2+} and SO_4^{2-}, (**c**) Fe^{3+} and CO_3^{2-}.

- Identify the locations of metals and nonmetals in the periodic table. (Section 2.5) *Related Exercises: 2.41, 2.42*
- Distinguish between molecular substances and ionic substances in terms of their composition. (Section 2.6 and 2.7) *Related Exercises: 2.61, 2.63, 2.65, 2.66*
- Distinguish between empirical formulas and molecular formulas. (Section 2.6) *Related Exercises: 2.45–2.47, 2.49*
- Describe how molecular formulas and structural formulas are used to represent the compositions of molecules. (Section 2.6) *Related Exercises: 2.53, 2.54*
- Explain how ions are formed by the gain or loss of electrons and use the periodic table to predict the charges of common ions. (Section 2.7) *Related Exercises: 2.56–2.58*
- Write the empirical formulas of ionic compounds, given the charges of their component ions. (Section 2.7) *Related Exercises: 2.59, 2.61–2.63*
- Write the name of an ionic compound given its chemical formula or write the chemical formula given its name. (Section 2.8) *Related Exercises: 2.71–2.74*
- Name or write chemical formulas for binary inorganic compounds and for acids. (Section 2.8) *Related Exercises: 2.75–2.78*
- Identify organic compounds and name simple alkanes and alcohols. (Section 2.9) *Related Exercises: 2.81, 2.83–2.85*

Key Equations

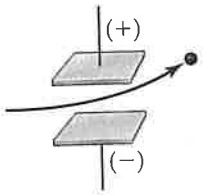
- Atomic weight $= \sum_{\text{over all isotopes}}[(\text{isotope mass}) \times (\text{fractional isotope abundance})]$ [2.1] Calculating atomic weight as a fractionally weighted average of isotopic masses.

Exercises

Visualizing Concepts

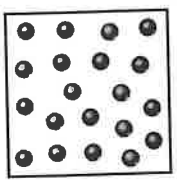
These exercises are intended to probe your understanding of key concepts rather than your ability to utilize formulas and perform calculations. Exercises with red numbers have answers in the back of the book.

2.1 A charged particle moves between two electrically charged plates, as shown here.

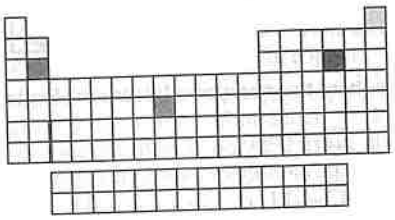


(a) What is the sign of the electrical charge on the particle? **(b)** As the charge on the plates is increased, would you expect the bending to increase, decrease, or stay the same? **(c)** As the mass of the particle is increased while the speed of the particles remains the same, would you expect the bending to increase, decrease, or stay the same? [Section 2.2]

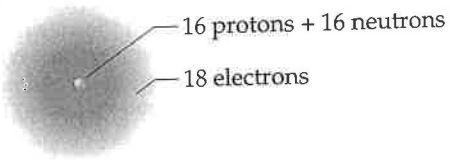
2.2 The following diagram is a representation of 20 atoms of a fictitious element, which we will call nevadium (Nv). The red spheres are ^{293}Nv, and the blue spheres are ^{295}Nv. **(a)** Assuming that this sample is a statistically representative sample of the element, calculate the percent abundance of each element. **(b)** If the mass of ^{293}Nv is 293.15 amu and that of ^{295}Nv is 295.15 amu, what is the atomic weight of Nv? [Section 2.4]



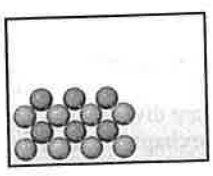
2.3 Four of the boxes in the following periodic table are colored. Which of these are metals and which are nonmetals? Which one is an alkaline earth metal? Which one is a noble gas? [Section 2.5]

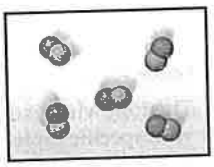


2.4 Does the following drawing represent a neutral atom or an ion? Write its complete chemical symbol, including mass number, atomic number, and net charge (if any). [Sections 2.3 and 2.7]



2.5 Which of the following diagrams most likely represents an ionic compound, and which represents a molecular one? Explain your choice. [Sections 2.6 and 2.7]

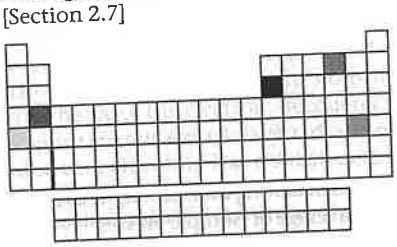

(i)


(ii)

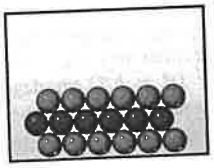
2.6 Write the chemical formula for the following compound. Is the compound ionic or molecular? Name the compound. [Sections 2.6 and 2.8]



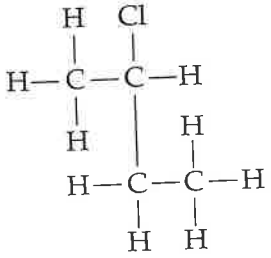
2.7 Five of the boxes in the following periodic table are colored. Predict the charge on the ion associated with each of these elements. [Section 2.7]

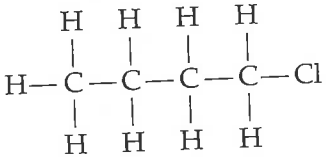


2.8 The following diagram represents an ionic compound in which the red spheres represent cations and the blue spheres represent anions. Which of the following formulas is consistent with the drawing? KBr, K_2SO_4, $Ca(NO_3)_2$, $Fe_2(SO_4)_3$. Name the compound. [Sections 2.7 and 2.8]

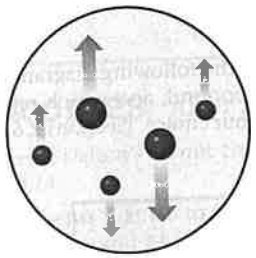


2.9 Are these two compounds isomers? Explain. [Section 2.9]





2.10 In the Millikan oil-drop experiment (see Figure 2.5), the tiny oil drops are observed through the viewing lens as rising, stationary, or falling, as shown here. (**a**) What causes their rate of fall to vary from their rate in the absence of an electric field? (**b**) Why do some drops move upward? [Section 2.2]



The following exercises are divided into sections that deal with specific topics in the chapter. The exercises are grouped in pairs, with the answers given in the back of the book to the odd-numbered exercises, as indicated by the red exercise numbers. Those exercises whose numbers appear in brackets are more challenging than the nonbracketed exercises.

The Atomic Theory of Matter and the Discovery of Atomic Structure (Sections 2.1 and 2.2)

2.11 A 1.0-g sample of carbon dioxide (CO_2) is fully decomposed into its elements, yielding 0.273 g of carbon and 0.727 g of oxygen. (**a**) What is the ratio of the mass of O to C? (**b**) If a sample of a different compound decomposes into 0.429 g of carbon and 0.571 g of oxygen, what is its ratio of the mass of O to C? (**c**) According to Dalton's atomic theory, what is the empirical formula of the second compound?

2.12 Hydrogen sulfide is composed of two elements: hydrogen and sulfur. In an experiment, 6.500 g of hydrogen sulfide is fully decomposed into its elements. (**a**) If 0.384 g of hydrogen is obtained in this experiment, how many grams of sulfur must be obtained? (**b**) What fundamental law does this experiment demonstrate?

2.13 A chemist finds that 30.82 g of nitrogen will react with 17.60, 35.20, 70.40, or 88.00 g of oxygen to form four different compounds. (**a**) Calculate the mass of oxygen per gram of nitrogen in each compound. (**b**) How do the numbers in part (a) support Dalton's atomic theory?

2.14 In a series of experiments, a chemist prepared three different compounds that contain only iodine and fluorine and determined the mass of each element in each compound:

Compound	Mass of Iodine (g)	Mass of Fluorine (g)
1	4.75	3.56
2	7.64	3.43
3	9.41	9.86

(**a**) Calculate the mass of fluorine per gram of iodine in each compound. (**b**) How do the numbers in part (a) support the atomic theory?

2.15 Which of the three subatomic particles was discovered first—the proton, neutron, or electron? Which one was discovered last?

2.16 An unknown particle is caused to move between two electrically charged plates, as illustrated in Figure 2.7. You hypothesize that the particle is a proton. (**a**) If your hypothesis is correct, would the particle be deflected in the same or opposite direction as the β rays? (**b**) Would it be deflected by a smaller or larger amount than the β rays?

2.17 What fraction of the α particles in Rutherford's gold foil experiment are scattered at large angles? Assume the gold foil is two layers thick, as shown in Figure 2.9, and that the approximate diameters of a gold atom and its nucleus are 2.7 Å and 1.0×10^{-4} Å, respectively. *Hint:* Calculate the cross sectional area occupied by the nucleus as a fraction of that occupied by the atom. Assume that the gold nuclei in each layer are offset from each other.

2.18 Millikan determined the charge on the electron by studying the static charges on oil drops falling in an electric field (Figure 2.5). A student carried out this experiment using several oil drops for her measurements and calculated the charges on the drops. She obtained the following data:

Droplet	Calculated Charge (C)
A	1.60×10^{-19}
B	3.15×10^{-19}
C	4.81×10^{-19}
D	6.31×10^{-19}

(a) What is the significance of the fact that the droplets carried different charges? (b) What conclusion can the student draw from these data regarding the charge of the electron? (c) What value (and to how many significant figures) should she report for the electronic charge?

The Modern View of Atomic Structure; Atomic Weights (Sections 2.3 and 2.4)

2.19 The radius of an atom of gold (Au) is about 1.35 Å. (a) Express this distance in nanometers (nm) and in picometers (pm). (b) How many gold atoms would have to be lined up to span 1.0 mm? (c) If the atom is assumed to be a sphere, what is the volume in cm^3 of a single Au atom?

2.20 An atom of rhodium (Rh) has a diameter of about 2.7×10^{-8} cm. (a) What is the radius of a rhodium atom in angstroms (Å) and in meters (m)? (b) How many Rh atoms would have to be placed side by side to span a distance of 6.0 μm? (c) If you assume that the Rh atom is a sphere, what is the volume in m^3 of a single atom?

2.21 Answer the following questions without referring to Table 2.1: (a) What are the main subatomic particles that make up the atom? (b) What is the relative charge (in multiples of the electronic charge) of each of the particles? (c) Which of the particles is the most massive? (d) Which is the least massive?

2.22 Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (a) The nucleus has most of the mass and comprises most of the volume of an atom. (b) Every atom of a given element has the same number of protons. (c) The number of electrons in an atom equals the number of neutrons in the atom. (d) The protons in the nucleus of the helium atom are held together by a force called the strong nuclear force.

2.23 Consider an atom of ^{10}B. (a) How many protons, neutrons, and electrons does this atom contain? (b) What is the symbol of the atom obtained by adding one proton to ^{10}B? (c) What is the symbol of the atom obtained by adding one neutron to ^{10}B? (d) Are either of the atoms obtained in parts (b) and (c) isotopes of ^{10}B? If so which one?

2.24 Consider an atom of ^{63}Cu. (a) How many protons, neutrons, and electrons does this atom contain? (b) What is the symbol of the ion obtained by removing two electrons from ^{63}Cu? (c) What is the symbol for the isotope of ^{63}Cu that possesses 36 neutrons?

2.25 (a) Define atomic number and mass number. (b) Which of these can vary without changing the identity of the element?

2.26 (a) Which two of the following are isotopes of the same element: $^{31}_{16}X$, $^{31}_{15}X$, $^{32}_{16}X$? (b) What is the identity of the element whose isotopes you have selected?

2.27 How many protons, neutrons, and electrons are in the following atoms? (a) ^{40}Ar, (b) ^{65}Zn, (c) ^{70}Ga, (d) ^{80}Br, (e) ^{184}W, (f) ^{243}Am.

2.28 Each of the following isotopes is used in medicine. Indicate the number of protons and neutrons in each isotope: (a) phosphorus-32, (b) chromium-51, (c) cobalt-60, (d) technetium-99, (e) iodine-131, (f) thallium-201.

2.29 Fill in the gaps in the following table, assuming each column represents a neutral atom.

Symbol	^{79}Br				
Protons		25			82
Neutrons		30	64		
Electrons			48	86	
Mass no.				222	207

2.30 Fill in the gaps in the following table, assuming each column represents a neutral atom.

Symbol	^{112}Cd				
Protons		38			92
Neutrons		58	49		
Electrons			38	36	
Mass no.				81	235

2.31 Write the correct symbol, with both superscript and subscript, for each of the following. Use the list of elements in the front inside cover as needed: (a) the isotope of platinum that contains 118 neutrons, (b) the isotope of krypton with mass number 84, (c) the isotope of arsenic with mass number 75, (d) the isotope of magnesium that has an equal number of protons and neutrons.

2.32 One way in which Earth's evolution as a planet can be understood is by measuring the amounts of certain isotopes in rocks. One quantity recently measured is the ratio of ^{129}Xe to ^{130}Xe in some minerals. In what way do these two isotopes differ from one another? In what respects are they the same?

2.33 (a) What isotope is used as the standard in establishing the atomic mass scale? (b) The atomic weight of boron is reported as 10.81, yet no atom of boron has the mass of 10.81 amu. Explain.

2.34 (a) What is the mass in amu of a carbon-12 atom? (b) Why is the atomic weight of carbon reported as 12.011 in the table of elements and the periodic table in the front inside cover of this text?

2.35 Only two isotopes of copper occur naturally: ^{63}Cu (atomic mass = 62.9296 amu; abundance 69.17%) and ^{65}Cu (atomic mass = 64.9278 amu; abundance 30.83%). Calculate the atomic weight (average atomic mass) of copper.

2.36 Rubidium has two naturally occurring isotopes, rubidium-85 (atomic mass = 84.9118 amu; abundance = 72.15%) and rubidium-87 (atomic mass = 86.9092 amu; abundance = 27.85%). Calculate the atomic weight of rubidium.

2.37 (a) Thomson's cathode-ray tube (Figure 2.4) and the mass spectrometer (Figure 2.11) both involve the use of electric or magnetic fields to deflect charged particles. What are the charged particles involved in each of these experiments? (b) What are the labels on the axes of a mass spectrum? (c) To measure the mass spectrum of an atom, the atom must first lose one or more electrons. Which would you expect to be deflected more by the same setting of the electric and magnetic fields, a Cl^+ or a Cl^{2+} ion?

2.38 Consider the mass spectrometer shown in Figure 2.11. Determine whether each of the following statements is true or false. If false, correct the statement to make it true: (a) The paths of neutral (uncharged) atoms are not affected by the magnet. (b) The height of each peak in the mass spectrum is inversely proportional to the mass of that isotope. (c) For a given element, the number of peaks in the spectrum is equal to the number of naturally occurring isotopes of that element.

2.39 Naturally occurring magnesium has the following isotopic abundances:

Isotope	Abundance (%)	Atomic mass (amu)
^{24}Mg	78.99	23.98504
^{25}Mg	10.00	24.98584
^{26}Mg	11.01	25.98259

(a) What is the average atomic mass of Mg? (b) Sketch the mass spectrum of Mg.

2.40 Mass spectrometry is more often applied to molecules than to atoms. We will see in Chapter 3 that the *molecular weight* of a molecule is the sum of the atomic weights of the atoms in the molecule. The mass spectrum of H_2 is taken under conditions that prevent decomposition into H atoms. The two naturally occurring isotopes of hydrogen are ^1H (atomic mass = 1.00783 amu; abundance 99.9885%) and ^2H (atomic mass = 2.01410 amu; abundance 0.0115%). (**a**) How many peaks will the mass spectrum have? (**b**) Give the relative atomic masses of each of these peaks. (**c**) Which peak will be the largest, and which the smallest?

The Periodic Table, Molecules, Molecular Compounds, Ions, and Ionic Compounds (Sections 2.5, 2.6, and 2.7)

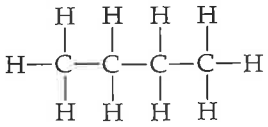
2.41 For each of the following elements, write its chemical symbol, locate it in the periodic table, give its atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) chromium, (**b**) helium, (**c**) phosphorus, (**d**) zinc, (**e**) magnesium, (**f**) bromine, (**g**) arsenic.

2.42 Locate each of the following elements in the periodic table; give its name and atomic number, and indicate whether it is a metal, metalloid, or nonmetal: (**a**) Li, (**b**) Sc, (**c**) Ge, (**d**) Yb, (**e**) Mn, (**f**) Sb, (**g**) Xe.

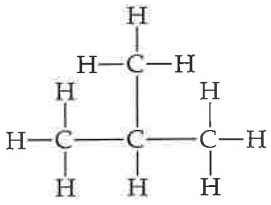
2.43 For each of the following elements, write its chemical symbol, determine the name of the group to which it belongs (Table 2.3), and indicate whether it is a metal, metalloid, or nonmetal: (**a**) potassium, (**b**) iodine, (**c**) magnesium, (**d**) argon, (**e**) sulfur.

2.44 The elements of group 4A show an interesting change in properties moving down the group. Give the name and chemical symbol of each element in the group and label it as a nonmetal, metalloid, or metal.

2.45 The structural formulas of the compounds *n*-butane and isobutane are shown below. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each. (**c**) Which formulas—empirical, molecular, or structural—allow you determine these are different compounds?

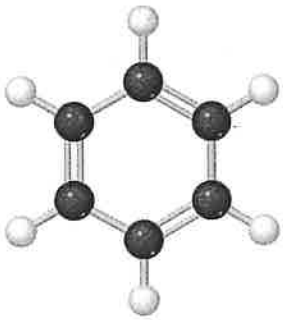


n-butane



isobutane

2.46 Ball-and-stick representations of benzene, a colorless liquid often used in organic chemistry reactions, and acetylene, a gas used as a fuel for high-temperature welding, are shown below. (**a**) Determine the molecular formula of each. (**b**) Determine the empirical formula of each.

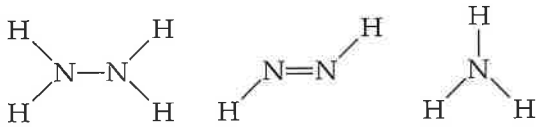


benzene



acetylene

2.47 What are the molecular and empirical formulas for each of the following compounds?



2.48 Two substances have the same molecular and empirical formulas. Does this mean that they must be the same compound?

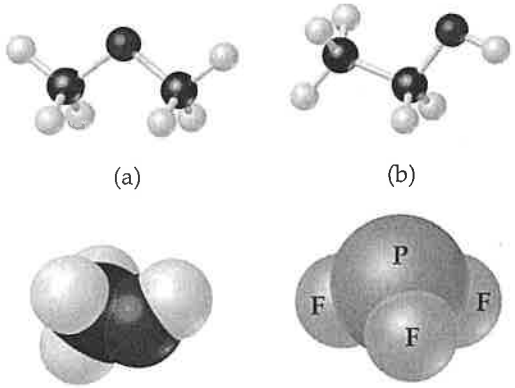
2.49 Write the empirical formula corresponding to each of the following molecular formulas: (**a**) Al_2Br_6, (**b**) C_8H_{10}, (**c**) $C_4H_8O_2$, (**d**) P_4O_{10}, (**e**) $C_6H_4Cl_2$, (**f**) $B_3N_3H_6$.

2.50 Determine the molecular and empirical formulas of the following: (**a**) the organic solvent *benzene*, which has six carbon atoms and six hydrogen atoms; (**b**) the compound *silicon tetrachloride*, which has a silicon atom and four chlorine atoms and is used in the manufacture of computer chips; (**c**) the reactive substance *diborane*, which has two boron atoms and six hydrogen atoms; (**d**) the sugar called *glucose*, which has six carbon atoms, twelve hydrogen atoms, and six oxygen atoms.

2.51 How many hydrogen atoms are in each of the following: (**a**) C_2H_5OH, (**b**) $Ca(C_2H_5COO)_2$, (**c**) $(NH_4)_3PO_4$?

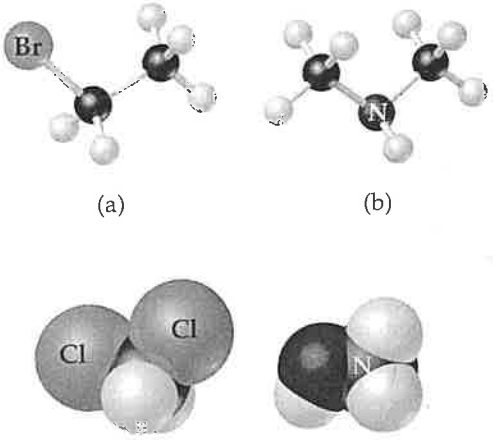
2.52 How many of the indicated atoms are represented by each chemical formula: (**a**) carbon atoms in $C_4H_9COOCH_3$, (**b**) oxygen atoms in $Ca(ClO_3)_2$, (**c**) hydrogen atoms in $(NH_4)_2HPO_4$?

2.53 Write the molecular and structural formulas for the compounds represented by the following molecular models:



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

2.54 Write the molecular and structural formulas for the compounds represented by the following models:



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

2.55 Fill in the gaps in the following table:

Symbol	$^{59}Co^{3+}$			
Protons		34	76	80
Neutrons		46	116	120
Electrons		36		78
Net charge			2+	

2.56 Fill in the gaps in the following table:

Symbol	$^{31}P^{3-}$			
Protons		34	50	
Neutrons		45	69	118
Electrons			46	76
Net charge		2−		3+

2.57 Each of the following elements is capable of forming an ion in chemical reactions. By referring to the periodic table, predict the charge of the most stable ion of each: (**a**) Mg, (**b**) Al, (**c**) K, (**d**) S, (**e**) F.

2.58 Using the periodic table, predict the charges of the ions of the following elements: (**a**) Ga, (**b**) Sr, (**c**) As, (**d**) Br, (**e**) Se.

2.59 Using the periodic table to guide you, predict the chemical formula and name of the compound formed by the following elements: (**a**) Ga and F, (**b**) Li and H, (**c**) Al and I, (**d**) K and S.

2.60 The most common charge associated with scandium in its compounds is 3+. Indicate the chemical formulas you would expect for compounds formed between scandium and (**a**) iodine, (**b**) sulfur, (**c**) nitrogen.

2.61 Predict the chemical formula for the ionic compound formed by (**a**) Ca^{2+} and Br^-, (**b**) K^+ and CO_3^{2-}, (**c**) Al^{3+} and CH_3COO^-, (**d**) NH_4^+ and SO_4^{2-}, (**e**) Mg^{2+} and PO_4^{3-}.

2.62 Predict the chemical formulas of the compounds formed by the following pairs of ions: (**a**) Cr^{3+} and Br^-, (**b**) Fe^{3+} and O^{2-}, (**c**) Hg_2^{2+} and CO_3^{2-}, (**d**) Ca^{2+} and ClO_3^-, (**e**) NH_4^+ and PO_4^{3-}.

2.63 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	K^+	NH_4^+	Mg^{2+}	Fe^{3+}
Cl^-	KCl			
OH^-				
CO_3^{2-}				
PO_4^{3-}				

2.64 Complete the table by filling in the formula for the ionic compound formed by each pair of cations and anions, as shown for the first pair.

Ion	Na^+	Ca^{2+}	Fe^{2+}	Al^{3+}
O^{2-}	Na_2O			
NO_3^-				
SO_4^{2-}				
AsO_4^{3-}				

2.65 Predict whether each of the following compounds is molecular or ionic: (**a**) B_2H_6, (**b**) CH_3OH, (**c**) $LiNO_3$, (**d**) Sc_2O_3, (**e**) CsBr, (**f**) NOCl, (**g**) NF_3, (**h**) Ag_2SO_4.

2.66 Which of the following are ionic, and which are molecular? (**a**) PF_5, (**b**) NaI, (**c**) SCl_2, (**d**) $Ca(NO_3)_2$, (**e**) $FeCl_3$, (**f**) LaP, (**g**) $CoCO_3$, (**h**) N_2O_4.

Naming Inorganic Compounds; Some Simple Organic Compounds (Sections 2.8 and 2.9)

2.67 Give the chemical formula for (**a**) chlorite ion, (**b**) chloride ion, (**c**) chlorate ion, (**d**) perchlorate ion, (**e**) hypochlorite ion.

2.68 Selenium, an element required nutritionally in trace quantities, forms compounds analogous to sulfur. Name the following ions: (**a**) SeO_4^{2-}, (**b**) Se^{2-}, (**c**) HSe^-, (**d**) $HSeO_3^-$.

2.69 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CaO, (**b**) Na_2SO_4, (**c**) $KClO_4$, (**d**) $Fe(NO_3)_2$, (**e**) $Cr(OH)_3$.

2.70 Give the names and charges of the cation and anion in each of the following compounds: (**a**) CuS, (**b**) Ag_2SO_4, (**c**) $Al(ClO_3)_3$, (**d**) $Co(OH)_2$, (**e**) $PbCO_3$.

2.71 Name the following ionic compounds: (**a**) Li_2O, (**b**) $FeCl_3$, (**c**) NaClO, (**d**) $CaSO_3$, (**e**) $Cu(OH)_2$, (**f**) $Fe(NO_3)_2$, (**g**) $Ca(CH_3COO)_2$, (**h**) $Cr_2(CO_3)_3$, (**i**) K_2CrO_4, (**j**) $(NH_4)_2SO_4$.

2.72 Name the following ionic compounds: (**a**) KCN, (**b**) $NaBrO_2$, (**c**) $Sr(OH)_2$, (**d**) CoTe, (**e**) $Fe_2(CO_3)_3$, (**f**) $Cr(NO_3)_3$, (**g**) $(NH_4)_2SO_3$, (**h**) NaH_2PO_4, (**i**) $KMnO_4$, (**j**) $Ag_2Cr_2O_7$.

2.73 Write the chemical formulas for the following compounds: (**a**) aluminum hydroxide, (**b**) potassium sulfate, (**c**) copper(I) oxide, (**d**) zinc nitrate, (**e**) mercury(II) bromide, (**f**) iron(III) carbonate, (**g**) sodium hypobromite.

2.74 Give the chemical formula for each of the following ionic compounds: (**a**) sodium phosphate, (**b**) zinc nitrate, (**c**) barium bromate, (**d**) iron(II) perchlorate, (**e**) cobalt(II) hydrogen carbonate, (**f**) chromium(III) acetate, (**g**) potassium dichromate.

2.75 Give the name or chemical formula, as appropriate, for each of the following acids: (**a**) $HBrO_3$, (**b**) HBr, (**c**) H_3PO_4, (**d**) hypochlorous acid, (**e**) iodic acid, (**f**) sulfurous acid.

2.76 Provide the name or chemical formula, as appropriate, for each of the following acids: (**a**) hydroiodic acid, (**b**) chloric acid, (**c**) nitrous acid, (**d**) H_2CO_3, (**e**) $HClO_4$, (**f**) CH_3COOH.

2.77 Give the name or chemical formula, as appropriate, for each of the following binary molecular substances: (**a**) SF_6, (**b**) IF_5, (**c**) XeO_3, (**d**) dinitrogen tetroxide, (**e**) hydrogen cyanide, (**f**) tetraphosphorus hexasulfide.

2.78 The oxides of nitrogen are very important components in urban air pollution. Name each of the following compounds: (**a**) N_2O, (**b**) NO, (**c**) NO_2, (**d**) N_2O_5, (**e**) N_2O_4.

2.79 Write the chemical formula for each substance mentioned in the following word descriptions (use the front inside cover to find the symbols for the elements you do not know). (**a**) Zinc carbonate can be heated to form zinc oxide and carbon dioxide. (**b**) On treatment with hydrofluoric acid, silicon dioxide forms silicon tetrafluoride and water. (**c**) Sulfur dioxide reacts with water to form sulfurous acid. (**d**) The substance phosphorus trihydride, commonly called phosphine, is a toxic gas. (**e**) Perchloric acid reacts with cadmium to form cadmium(II) perchlorate. (**f**) Vanadium(III) bromide is a colored solid.

2.80 Assume that you encounter the following sentences in your reading. What is the chemical formula for each substance mentioned? (**a**) Sodium hydrogen carbonate is used as a deodorant. (**b**) Calcium hypochlorite is used in some bleaching solutions. (**c**) Hydrogen cyanide is a very poisonous gas. (**d**) Magnesium hydroxide is used as a cathartic. (**e**) Tin(II) fluoride has been used as a fluoride additive in toothpastes. (**f**) When cadmium sulfide is treated with sulfuric acid, fumes of hydrogen sulfide are given off.

2.81 (**a**) What is a hydrocarbon? (**b**) Pentane is the alkane with a chain of five carbon atoms. Write a structural formula for this compound and determine its molecular and empirical formulas.

ANSWERS TO SELECTED PRACTICE EXERCISES

Chapter 1

Sample Exercise 1.1
Practice Exercise 2: It is a compound because it has constant composition and can be separated into several elements.

Sample Exercise 1.2
Practice Exercise 2: (**a**) 10^{12} pm, (**b**) 6.0 km, (**c**) 4.22×10^{-3} g, (**d**) 0.00422 g.

Sample Exercise 1.3
Practice Exercise 2: (**a**) 261.7 K, (**b**) 11.3 °F

Sample Exercise 1.4
Practice Exercise 2: (**a**) 8.96 g/cm^3, (**b**) 19.0 mL, (**c**) 340 g.

Sample Exercise 1.5
Practice Exercise 2: 2.29×10^6 J

Sample Exercise 1.6
Practice Exercise 2: No. The number of feet in a mile is a defined quantity and is therefore exact, but the distance represented by one foot is not exact, although it is known to high accuracy.

Sample Exercise 1.7
Practice Exercise 2: (**a**) four, (**b**) two, (**c**) three.

Sample Exercise 1.8
Practice Exercise 2: 9.52 m/s (three significant figures).

Sample Exercise 1.9
Practice Exercise 2: No. Even though the mass of the gas would then be known to four significant figures, the volume of the container would still be known to only three.

Sample Exercise 1.10
Practice Exercise 2: 804.7 km

Sample Exercise 1.11
Practice Exercise 2: 12 km/L.

Sample Exercise 1.12
Practice Exercise 2: 1.2×10^4 ft.

Sample Exercise 1.13
Practice Exercise 2: 832 g

Chapter 2

Sample Exercise 2.1
Practice Exercise 2: (**a**) 154 pm, (**b**) 1.3×10^6 C atoms

Sample Exercise 2.2
Practice Exercise 2: (**a**) 56 protons, 56 electrons, and 82 neutrons, (**b**) 15 protons, 15 electrons, and 16 neutrons.

Sample Exercise 2.3
Practice Exercise 2: $^{208}_{82}Pb$

Sample Exercise 2.4
Practice Exercise 2: 28.09 amu

Sample Exercise 2.5
Practice Exercise 2: Na, atomic number 11, is a metal; Br, atomic number 35, is a nonmetal.

Sample Exercise 2.6
Practice Exercise 2: B_5H_7

Sample Exercise 2.7
Practice Exercise 2: 34 protons, 45 neutrons, and 36 electrons

Sample Exercise 2.8
Practice Exercise 2: (**a**) 3+, (**b**) 1−

Sample Exercise 2.9
Practice Exercise 2: (**a**) Rb is from group 1, and readily loses one electron to attain the electron configuration of the nearest noble gas element, Kr. (**b**) Nitrogen and the halogens are all nonmetallic elements, which form molecular compounds with one another. (**c**) Krypton, Kr, is a noble gas element and is chemically inactive except under special conditions. (**d**) Na and K are both from group 1 and adjacent to one another in the periodic table. They would be expected to behave very similarly. (**e**) Calcium is an active metal and readily loses two electrons to attain the noble gas configuration of Ar.

Sample Exercise 2.10
Practice Exercise 2: (**a**) Na_3PO_4, (**b**) $ZnSO_4$, (**c**) $Fe_2(CO_3)_3$

Sample Exercise 2.11
Practice Exercise 2: BrO^- and BrO_2^-

Sample Exercise 2.12
Practice Exercise 2: (**a**) ammonium bromide, (**b**) chromium(III) oxide, (**c**) cobalt(II) nitrate

Sample Exercise 2.13
Practice Exercise 2: (**a**) HBr, (**b**) H_2CO_3

Sample Exercise 2.14
Practice Exercise 2: (**a**) $SiBr_4$, (**b**) S_2Cl_2, (**c**) P_2O_6.

Sample Exercise 2.15
Practice Exercise 2: No, they are not isomers because they have different molecular formulas. Butane is C_4H_{10}, whereas cyclobutane is C_4H_8.

Chapter 3

Sample Exercise 3.1
Practice Exercise 2: (**a**) $C_2H_4 + 3\,O_2 \longrightarrow 2\,CO_2 + 2\,H_2O$.
(**b**) Nine O_2 molecules

Sample Exercise 3.2
Practice Exercise 2: (**a**) 4, 3, 2; (**b**) 2, 6, 2, 3; (**c**) 1, 2, 1, 1, 1

Sample Exercise 3.3
Practice Exercise 2: (**a**) $HgS(s) \longrightarrow Hg(l) + S(s)$, or $\frac{1}{8}S_8(s)$
(**b**) $4\,Al(s) + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$

Sample Exercise 3.4
Practice Exercise 2: $C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

Sample Exercise 3.5
Practice Exercise 2: (**a**) 78.0 amu, (**b**) 32.0 amu, (**c**) 211.0 amu

Sample Exercise 3.6
Practice Exercise 2: 16.1%

ANSWERS TO SELECTED EXERCISES

Chapter 1

1.1 (**a**) Pure element: i (**b**) mixture of elements: v, vi (**c**) pure compound: iv (**d**) mixture of an element and a compound: ii, iii **1.3** (**a**) Homogeneous mixture (**b**) Yes, brass is a solution. **1.5** Filtration **1.7** (**a**) The aluminum sphere is lightest, then nickel, then silver. (**b**) The platinum cube is smallest, then gold, then lead. **1.9** (**a**) 7.5 cm; two significant figures (sig figs) (**b**) 72 mi/hr (inner scale, two significant figures) or 115 km/hr (outer scale, three significant figures) **1.11** 464 jelly beans. The mass of an average bean has 2 decimal places and 3 sig figs. The number of beans then has 3 sig figs, by the rules for multiplication and division. **1.13** (**a**) Heterogeneous mixture (**b**) homogeneous mixture (heterogeneous if there are undissolved particles) (**c**) pure substance (**d**) pure substance. **1.15** (**a**) S (**b**) Au (**c**) K (**d**) Cl (**e**) Cu (**f**) uranium (**g**) nickel (**h**) sodium (**i**) aluminum (**j**) silicon **1.17** C is a compound; it contains both carbon and oxygen. A is a compound; it contains at least carbon and oxygen. B is not defined by the data given; it is probably also a compound because few elements exist as white solids. **1.19** Physical properties: silvery white; lustrous; melting point = 649 °C, boiling point = 1105 °C; density at 20 °C = 1.738 g/cm^3; pounded into sheets; drawn into wires; good conductor. Chemical properties: burns in air; reacts with Cl_2. **1.21** (**a**) Chemical (**b**) physical (**c**) physical (**d**) chemical (**e**) chemical **1.23** Distillation **1.25** (**a**) 1.9×10^5 J (**b**) 4.6×10^4 cal (**c**) The "lost" kinetic energy is converted mostly to heat. **1.27** (**a**) Kinetic energy (**b**) Potential energy decreases. **1.29** (**a**) $E_k = 11$ J; v = 6.6 m/s **1.31** (**a**) 1×10^{-1} (**b**) 1×10^{-2} (**c**) 1×10^{-15} (**d**) 1×10^{-6} (**e**) 1×10^{6} (**f**) 1×10^{3} (**g**) 1×10^{-9} (**h**) 1×10^{-3} (**i**) 1×10^{-12} **1.33** (**a**) 22 °C (**b**) 422.1 °F (**c**) 506 K (**d**) 107 °F (**e**) 1600 K (**f**) −459.67 °F **1.35** (**a**) 1.62 g/mL. Tetrachloroethylene, 1.62 g/mL, is more dense than water, 1.00 g/mL; tetrachloroethylene will sink rather than float on water. (**b**) 11.7 g **1.37** (**a**) Calculated density = 0.86 g/mL. The substance is probably toluene, density = 0.866 g/mL. (**b**) 40.4 mL ethylene glycol (**c**) Yes, the graduated cylinder in Figure 1.21 will provide the appropriate level of accuracy. (**d**) 1.11×10^3 g nickel **1.39** 36 Pg **1.41** 1054 J/Btu **1.43** Exact: (**b**), (**d**), and (**f**) **1.45** (**a**) 3 (**b**) 2 (**c**) 5 (**d**) 3 (**e**) 5 (**f**) 1 **1.47** (**a**) 1.025×10^2 (**b**) 6.570×10^2 (**c**) 8.543×10^{-3} (**d**) 2.579×10^{-4} (**e**) -3.572×10^{-2} **1.49** (**a**) 17.00 (**b**) 812.0 (**c**) 8.23×10^3 (**d**) 8.69×10^{-2} **1.51** 5 significant figures

1.53 (**a**) $\frac{1 \times 10^{-3}\ \text{m}}{1\text{mm}} \times \frac{1\text{nm}}{1 \times 10^{-9}\ \text{m}}$ (**b**) $\frac{1 \times 10^{-3}\ \text{g}}{1\text{mg}} \times \frac{1\ \text{kg}}{1000\ \text{g}}$

(**c**) $\frac{1000\ \text{m}}{1\text{km}} \times \frac{1\text{cm}}{1 \times 10^{-2}\ \text{m}} \times \frac{1\ \text{in.}}{2.54\ \text{cm}} \times \frac{1\text{ft}}{12\ \text{in.}}$ (**d**) $\frac{(2.54)^3\ \text{cm}^3}{1^3\ \text{in.}^3}$

1.55 (**a**) 54.7 km/hr (**b**) 1.3×10^3 gal (**c**) 46.0 m (**d**) 0.984 in/hr **1.57** (**a**) 4.32×10^5 s (**b**) 88.5 m (**c**) $0.499/L (**d**) 46.6km/hr (**e**) 1.420 L/s (**f**) 707.9 cm^3 **1.59** (**a**) 1.2×10^2 L (**b**) 5×10^2 mg (**c**) 19.9mi/gal (2×10^1 mi/gal for 1 significant figure) (**d**) 1.81 kg **1.61** 64 kg air **1.63** $6 \times 10^4 **1.67** 8.47 g O **1.69** (**a**) Set I, 22.51; set II, 22.61. Based on the average, set I is more accurate. (**b**) The average deviation for both set I and set II is 0.02. The two sets display the same precision. **1.71** (**a**) Volume (**b**) area (**c**) volume (**d**) density (**e**) time (**f**) length (**g**) temperature **1.74** Substances (**c**), (**d**), (**e**), (**g**) and (**h**) are pure or nearly pure. **1.75** (**a**) 1.13×10^5 quarters (**b**) 6.41×10^5 g (**c**) $2.83 \times 10^4 **(d**) 5.74×10^8 stacks **1.79** The most dense liquid, Hg, will sink; the least dense, cyclohexane, will float; H_2O will be in the middle. **1.82** density of solid = 1.63 g/mL **1.85** (**a**) Density of peat = 0.13 g/cm^3, density of soil = 2.5g/cm^3. It is not correct to say that peat is "lighter" than topsoil. Volumes must be specified in order to compare masses. (**b**) Buy 16 bags of peat (more than 15 are needed). (Results to 1 significant figure are not meaningful.) **1.89** The inner diameter of the tube is 1.71 cm. **1.91** The separation with distinctly separated red and blue spots is more successful. To quantify the characteristics of the separation, calculate a reference value for each spot: distance traveled by spot/distance traveled by solvent. If the values for the two spots are fairly different, the separation is successful. **1.94** (**a**) Volume = 0.050mL (**b**) surface area = 12.4 m^2 (**c**) 99.99% of the mercury was removed. (**d**) The spongy material weighs 17.7 mg after exposure to mercury.

Chapter 2

2.1 (**a**) (−) (**b**) increase (**c**) decrease **2.4** The particle is an ion. $^{32}_{16}S^{2-}$ **2.6** Formula: IF_5; name: iodine pentafluoride; the compound is molecular. **2.8** Only $Ca(NO_3)_2$, calcium nitrate, is consistent with the diagram. **2.10** (**a**) In the presence of an electric field, there is electrostatic attraction between the negatively charged oil drops and the positively charged plate as well as electrostatic repulsion between the negatively charged oil drops and the negatively charged plate. These electrostatic forces oppose the force of gravity and change the rate of fall of the drops. (**b**) Each individual drop has a different number of electrons associated with it. If the combined electrostatic forces are greater than the force of gravity, the drop moves up. **2.11** (**a**) mass O/mass C = 2.66 (**b**) mass O/mass C = 1.33 (**c**) CO **2.13** (**a**) 0.5711 g O/1 g N; 1.142 g O/1 g N; 2.284 g O/1 g N; 2.855 g O/1 g N (**b**) The numbers in part (**a**) obey the *law of multiple proportions*. Multiple proportions arise because atoms are the indivisible entities combining, as stated in Dalton's atomic theory. **2.15** Electrons, in the form of cathode rays, were discovered first. Neutrons were discovered last. **2.17** The fraction of alpha particles deflected at a large angle is 2.7×10^{-9}. That is, 1 out of approximately 365 million alpha particles is deflected at a large angle. **2.19** (**a**) 0.135 nm; 1.35 × 10^2 or 135 pm (**b**) 3.70×10^6 Au atoms (**c**) 1.03×10^{-23} cm^3 **2.21** (**a**) Proton, neutron, electron (**b**) proton = 1+, neutron = 0, electron = 1− (**c**) The neutron is most massive. (The neutron and proton have very similar masses.) (**d**) The electron is least massive. **2.23** (**a**) 5 protons, 5 neutrons, 5 electrons (**b**) $^{11}_{6}C$ (**c**) $^{11}_{5}B$ (**d**) The atom in part (**c**) is an isotope of ^{10}B. **2.25** (**a**) Atomic number is the number of protons in the nucleus of an atom. Mass number is the total number of nuclear particles, protons plus neutrons, in an atom. (**b**) mass number **2.27** (**a**) ^{40}Ar: 18 p, 22 n, 18 e (**b**) ^{65}Zn: 30 p, 35 n, 30 e (**c**) ^{70}Ga: 31 p, 39 n, 31 e (**d**) ^{80}Br: 35 p, 45 n, 35 e (**e**) ^{184}W: 74 p, 110 n, 74 e (**f**) ^{243}Am: 95 p, 148 n, 95e

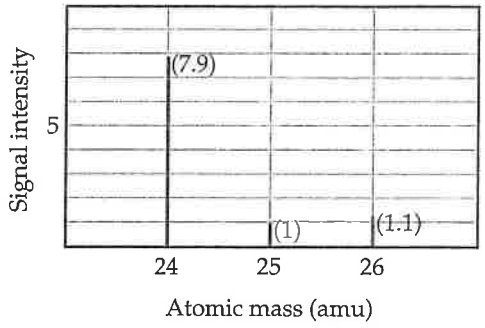
2.29

Symbol	^{79}Br	^{55}Mn	^{112}Cd	^{222}Rn	^{207}Pb
Protons	35	25	48	86	82
Neutrons	44	30	64	136	125
Electrons	35	25	48	86	82
Mass no.	79	55	112	222	207

2.31 (**a**) $^{196}_{78}Pt$ (**b**) $^{84}_{36}Kr$ (**c**) $^{75}_{33}As$ (**d**) $^{24}_{12}Mg$ **2.33** (**a**) $^{12}_{6}C$ (**b**) Atomic weights are average atomic masses, the sum of the mass of each

naturally occurring isotope of an element times its fractional abundance. Each B atom will have the mass of one of the naturally occurring isotopes, while the "atomic weight" is an average value. **2.35** 63.55 amu **2.37** (**a**) In Thomson's cathode ray tube, the charged particles are electrons. In a mass spectrometer, the charged particles are positively charged ions (cations). (**b**) The *x*-axis label is atomic mass (or particle mass), and the *y*-axis label is signal intensity. (**c**) The Cl^{2+} ion will be deflected more. **2.39** (**a**) average atomic mass = 24.31 amu

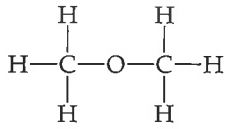
(**b**)



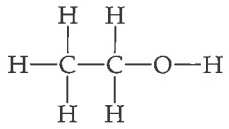
2.41 (**a**) Cr, 24 (metal) (**b**) He, 2 (nonmetal) (**c**) P, 15 (nonmetal) (**d**) Zn, 30 (metal) (**e**) Mg, 12 (metal) (**f**) Br, 35 (nonmetal) (**g**) As, 33 (metalloid) **2.43** (**a**) K, alkali metals (metal) (**b**) I, halogens (nonmetal) (**c**) Mg, alkaline earth metals (metal) (**d**) Ar, noble gases (nonmetal) (**e**) S, chalcogens (nonmetal) **2.45** (**a**) C_4H_{10} is the molecular formula for both compounds. (**b**) C_2H_5 is the empirical formula for both compounds. (**c**) structural **2.47** From left to right: molecular, N_2H_4, empirical, NH_2; molecular, N_2H_2, empirical, NH, molecular and empirical, NH_3 **2.49** (**a**) $AlBr_3$ (**b**) C_4H_5 (**c**) C_2H_4O (**d**) P_2O_5 (**e**) C_3H_2Cl (**f**) BNH_2 **2.51** (**a**) 6 (**b**) 10 (**c**) 12

2.53

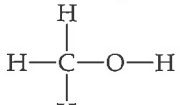
(**a**) C_2H_6O,



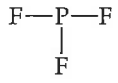
(**b**) C_2H_6O,



(**c**) CH_4O,



(**d**) PF_3,



2.55

Symbol	$^{59}Co^{3+}$	$^{80}Se^{2-}$	$^{192}Os^{2+}$	$^{200}Hg^{2+}$
Protons	27	34	76	80
Neutrons	32	46	116	120
Electrons	24	36	74	78
Net Charge	3+	2−	2+	2+

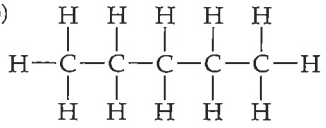
2.57 (**a**) Mg^{2+} (**b**) Al^{3+} (**c**) K^+ (**d**) S^{2-} (**e**) F^- **2.59** (**a**) GaF_3, gallium(III) fluoride (**b**) LiH, lithium hydride (**c**) AlI_3, aluminum iodide (**d**) K_2S, potassium sulfide **2.61** (**a**) $CaBr_2$ (**b**) K_2CO_3 (**c**) $Al(CH_3COO)_3$ (**d**) $(NH_4)_2SO_4$ (**e**) $Mg_3(PO_4)_2$

2.63

Ion	K^+	NH_4^+	Mg^{2+}	Fe^{3+}
Cl^-	KCl	NH_4Cl	$MgCl_2$	$FeCl_3$
OH^-	KOH	NH_4OH	$Mg(OH)_2$	$Fe(OH)_3$
CO_3^{2-}	K_2CO_3	$(NH_4)_2CO_3$	$MgCO_3$	$Fe_2(CO_3)_3$
PO_4^{3-}	K_3PO_4	$(NH_4)_3PO_4$	$Mg_3(PO_4)_2$	$FePO_4$

2.65 Molecular: (**a**) B_2H_6 (**b**) CH_3OH (**f**) NOCl (**g**) NF_3. Ionic: (**c**) $LiNO_3$ (**d**) Sc_2O_3 (**e**) CsBr (**f**) Ag_2SO_4 **2.67** (**a**) ClO_2^- (**b**) Cl^- (**c**) ClO_3^- (**d**) ClO_4^- (**e**) ClO^- **2.69** (**a**) calcium, 2+; oxide, 2− (**b**) sodium, 1+; sulfate, 2− (**c**) potassium, 1+; perchlorate, 1− (**d**) iron, 2+, nitrate, 1− (**e**) chromium, 3+; hydroxide, 1− **2.71** (**a**) lithium oxide (**b**) iron(III) chloride (ferric chloride) (**c**) sodium hypochlorite (**d**) calcium sulfite (**e**) copper(II) hydroxide (cupric hydroxide) (**f**) iron(II) nitrate (ferrous nitrate) (**g**) calcium acetate (**h**) chromium(III) carbonate (chromic carbonate) (**i**) potassium chromate (**j**) ammonium sulfate **2.73** (**a**) $Al(OH)_3$ (**b**) K_2SO_4 (**c**) Cu_2O (**d**) $Zn(NO_3)_2$ (**e**) $HgBr_2$ (**f**) $Fe_2(CO_3)_3$ (**g**) NaBrO **2.75** (**a**) Bromic acid (**b**) hydrobromic acid (**c**) phosphoric acid (**d**) HClO (**e**) HIO_3 (**f**) H_2SO_3 **2.77** (**a**) Sulfur hexafluoride (**b**) iodine pentafluoride (**c**) xenon trioxide (**d**) N_2O_4 (**e**) HCN (**f**) P_4S_6 **2.79** (**a**) $ZnCO_3$, ZnO, CO_2 (**b**) HF, SiO_2, SiF_4, H_2O (**c**) SO_2, H_2O, H_2SO_3 (**d**) PH_3 (**e**) $HClO_4$, Cd, $Cd(ClO_4)_2$ (**f**) VBr_3 **2.81** (**a**) A hydrocarbon is a compound composed of the elements hydrogen and carbon only.

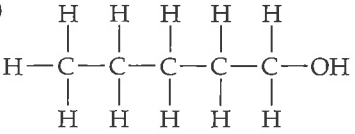
(**b**)



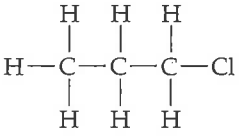
molecular and empirical formulas, C_5H_{12}

2.83 (**a**) A functional group is a group of specific atoms that are constant from one molecule to the next. (**b**) —OH

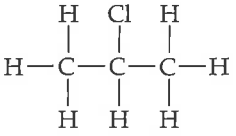
(**c**)



2.85 (a, b)



1-chloropropane



2-chloropropane

2.88 (**a**) 2 protons, 1 neutron, 2 electrons (**b**) tritium, ^3H, is more massive. (**c**) A precision of 1×10^{-27} g would be required to differentiate between $^3H^+$ and $^3He^+$. **2.90** Arrangement A, 4.1×10^{14} atoms/cm^2 (**b**) Arrangement B, 4.7×10^{14} atoms/cm^2 (**c**) The ratio of atoms going from arrangement B to arrangement A is 1.2 to 1. In three dimensions, arrangement B leads to a greater density for Rb metal. **2.94** (**a**) $^{16}_{8}O$, $^{17}_{8}O$, $^{18}_{8}O$ (**b**) All isotopes are atoms of the same element, oxygen, with the same atomic number, 8 protons in the nucleus and 8 electrons. We expect their electron arrangements to be the same and their chemical properties to be very similar. Each has a different number of neutrons, a different mass number, and a different atomic mass. **2.96** (**a**) $^{69}_{31}Ga$, 31 protons, 38 neutrons; $^{71}_{31}Ga$, 31 protons, 40 neutrons (**b**) $^{69}_{31}Ga$, 60.3%, $^{71}_{31}Ga$, 39.7%. **2.99** (**a**) 5 significant figures (**b**) An electron is 0.05444% of the mass of an ^1H atom.