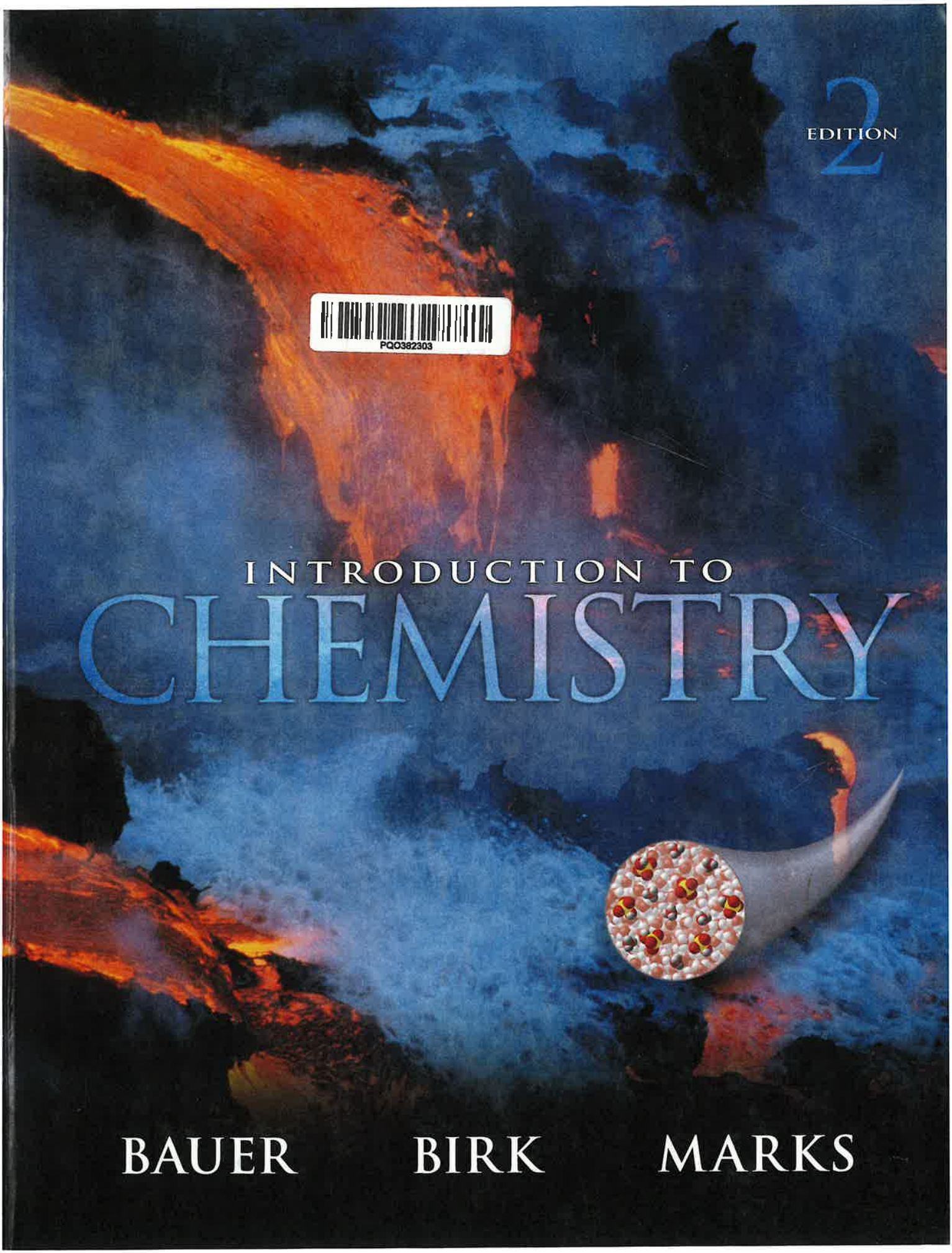

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Higher Education

INTRODUCTION TO CHEMISTRY: A CONCEPTUAL APPROACH, SECOND EDITION

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CHAPTER

1

Matter and Energy

1.1 Matter and Its Classification

1.2 Physical and Chemical Changes and Properties of Matter

1.3 Energy and Energy Changes

1.4 Scientific Inquiry

Summary

Math Toolbox 1.1 Scientific Notation

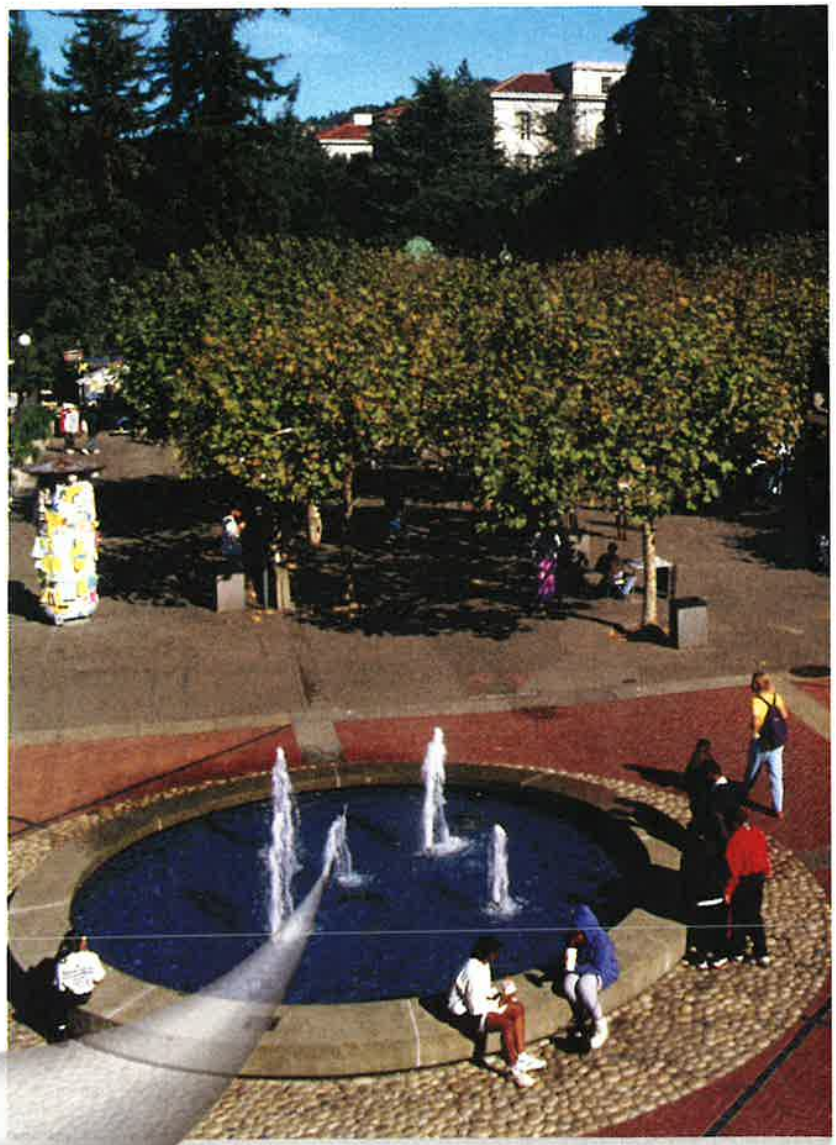
Math Toolbox 1.2 Significant Figures

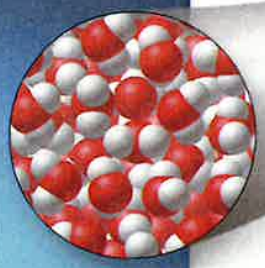
Math Toolbox 1.3 Units and Conversions

Key Relationships

Key Terms

Questions and Problems





Anna and Bill are enrolled in an introductory chemistry course. For their first assignment, the professor has asked them to walk around campus, locate objects that have something to do with chemistry, and classify the things they find according to characteristics of structure and form.

Anna and Bill begin their trek at the bookstore. They spot a fountain, a large metallic sculpture, a building construction site, and festive balloons decorating the front of the store. They notice water splashing in the fountain and coins that have collected at the bottom. The metallic sculpture has a unique color and texture. At the building construction site they notice murals painted on the wooden safety barricade. Through a hole in the fence, they see a construction worker doing some welding.

Bill and Anna make a list of the things that attracted their attention and start trying to classify them. Inspecting the fountain, they notice that it appears to be composed of pebbles embedded in cement. As water circulates in the fountain, it travels in waves on the water's surface. The coins in the fountain, mostly pennies, vary in their shininess. Some look new, with their copper color gleaming in the bright sunshine. Others look dingy, brown, and old. The metal sculpture has a unique, modern design, but it's showing signs of age. A layer of rust covers its entire surface. Anna and Bill decide to classify the sculpture as a metal, like the coins in the fountain. They also conclude that the water, pebbles, and concrete in the fountain are not metals.

As they approach the construction site, Anna and Bill examine the painted mural. Through the peephole in the mural, they see gravel, cinder blocks, metallic tubes for ductwork, steel beams, and copper pipe. They add more nonmetals and metals to their list. A welder is joining two pieces of metal. Sparks are flying everywhere. Anna and Bill wonder what is in the sparks. Since the sparks are so small and vanish so rapidly, they don't know how to classify them.

As they continue their walk, they pass the intramural fields and the gym where they see students using tennis rackets, baseball bats, bicycles, and weight belts. They wonder how they will classify these items. For lunch, Bill and Anna buy pizza. They sip soft drinks from aluminum cans. They settle on a bench to enjoy their lunch in the sunshine and watch students playing volleyball in a sandpit. As they put on their sunscreen, they wonder how they might classify sunlight. After lunch, they hurry off to an afternoon class. On the way, they notice a variety of vehicles on campus. Some are gasoline-powered cars and buses, but others have signs on them saying they operate on alternative fuels. Trucks lumber by, exhaust fumes spewing from their tailpipes. Bill and Anna feel the hoods of parked cars. Some are still warm from their engine's heat.

How are Bill's and Anna's observations related to chemistry? What characteristics have they identified that they can use for classification purposes? They have started their classification with metals and nonmetals. What other categories should they devise?

Now it's your turn. Make a list of things relevant to chemistry in the location where you are reading this. How will you classify the things on your list? What characteristics will you use to organize the items into categories? Most important, why bother to classify things at all?

In this chapter we will explore some answers to these questions. As you learn what chemistry is, you'll begin to develop explanations for how substances look, change, and behave.

Questions for Consideration

1.1 What characteristics distinguish different types of matter?
1.2 What are some properties of matter?
1.3 What is energy and how does it differ from matter?
1.4 What approaches do scientists use to answer these and other questions?

This icon refers to a Math Toolbox that provides more detail and practice.

Math Tools Used in This Chapter

Scientific Notation (Math Toolbox 1.1)
Significant Figures (Math Toolbox 1.2)
Units and Conversions (Math Toolbox 1.3)

1.1 | MATTER AND ITS CLASSIFICATION

All the *things* that Anna and Bill observed on campus are examples of matter. The fountain, the metal sculpture, the construction site, the balloons outside the bookstore, the exhaust fumes from buses, the pizza they had for lunch, even Bill and Anna themselves—all are matter. **Matter** is anything that occupies space and has mass. **Mass** is a measure of the quantity of matter. The interaction of mass with gravity creates weight, which can be measured on a scale or balance.

Some of Bill's and Anna's observations, however, were not of matter. Sunlight, the light from welding, and the heat of automobile engines are not matter. They do not occupy space, and they have no mass. They are forms of energy. *Energy* is the capacity to move an object or to transfer heat. We'll discuss energy in Section 1.3, but for now, let's focus on matter.

All of Anna's and Bill's observations are relevant to chemistry, because chemistry is the study of matter and energy. Since the entire physical world is matter and energy, chemistry would be an overwhelming subject of study if we did not classify phenomena in manageable ways. Anna and Bill used characteristics like shininess and hardness when they decided some materials were metals and others were not. Let's explore some other characteristics that can be used to classify matter.

Composition of Matter

One way to classify matter is by its chemical composition. Some types of matter always have the same chemical composition, no matter what their origin. Such matter is called a **pure substance** or more briefly, a *substance*. A pure substance has the same composition throughout and from sample to sample. It cannot be separated into components by physical means.

Some pure substances can be observed. For example, the aluminum in Anna's soda can is pure. It is not combined with any other substances, although it is coated with plastic and paint. Consider also the sandpit where Bill and Anna watched the volleyball game. The sand is not a pure substance, but if we removed all the dirt, minerals, and other contaminants, it would be the pure substance, silica, which is one kind of sand (Figure 1.1). Grains of silica differ in size, but they all have the same chemical composition, which can be determined in the laboratory.



FIGURE 1.1 Sand is composed of a mineral, silica. It contains the elements silicon and oxygen in specific proportions.

In contrast to pure substances, other materials are mixtures. A **mixture** consists of two or more pure substances and may vary in composition. The fountain, for example, is made from a mixture of gravel, concrete, and pebbles. Even the water in the fountain is not a pure substance since small amounts of gases and minerals are dissolved in it. Like sand, however, it could be made pure if all the other substances were removed.

Are there any things where you are now that might be pure substances? Actually, pure substances are rare in our world. Most things are mixtures of some kind. Pure substances are found most often in laboratories where they are used to determine the properties and behavior of matter under controlled conditions.

Elements All matter consists of pure substances or mixtures of substances. Pure substances, in turn, are of two types: elements and compounds. An **element** is a substance that cannot be broken down into simpler substances *even by a chemical reaction.* For example, suppose we first purified the water in a fountain to remove

contaminants. Then we used a chemical process called *electrolysis* to separate it into its component elements. Water can be broken down by chemical means into hydrogen and oxygen, as shown in Figure 1.2, so water is not an element. The hydrogen and oxygen, however, are elements. We cannot break them down into any simpler substances using heat, light, electricity, or any chemical process. We can convert them into more complex substances, but not into simpler ones.

Elements are the building blocks of all matter. Of the 111 elements that have been given names, 83 can be found in natural substances and in sufficient quantity to isolate. The many examples of matter that we use, see, and read about are all built up of different elements in different combinations. The elements that are not isolated from natural sources on Earth have been synthesized by scientists. Some are so unstable that they have only a fleeting existence, including those that have not yet been formally named. To classify elements, chemists use a *periodic table,* like that shown in Figure 1.3. The elements in each column, called *groups* or *families* of elements in the periodic table, share similar characteristics, or *properties.*

Elements are generally classified into two main categories: metals and nonmetals. Generally, a **metal** can be distinguished from a **nonmetal** by its luster (shininess) and ability to conduct electricity (electrical conductivity). Copper, aluminum, iron, and other metals are good conductors of electricity. Nonmetal elements, such as carbon (in the form of diamond), chlorine, and sulfur, normally are not. Note the

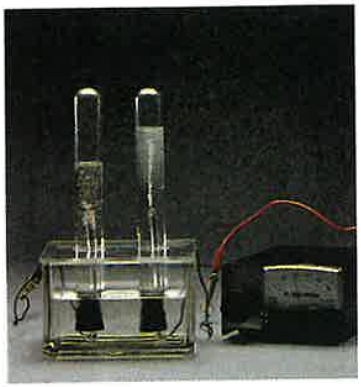


FIGURE 1.2 When electric current is passed through water, the water decomposes into the elements hydrogen and oxygen. The hydrogen (left) and oxygen (right) can be seen bubbling to the tops of the tubes.

Metals (main-group)
Metals (transition)
Metals (inner-transition)
Metalloids
Nonmetals

Period	MAIN-GROUP ELEMENTS IA (1)	IIA (2)	TRANSITION ELEMENTS IIIB (3)	IVB (4)	VB (5)	VIB (6)	VIIB (7)	VIIIB (8)	VIIIB (9)	VIIIB (10)	IB (11)	IIB (12)	MAIN-GROUP ELEMENTS IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.41	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (263)	105 Db (262)	106 Sg (266)	107 Bh (267)	108 Hs (277)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 (285)	113 (284)	114 (289)	115 (289)	116 (292)		

INNER-TRANSITION ELEMENTS

6	Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
7	Actinides	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

FIGURE 1.3 The periodic table organizes the known elements according to their properties. The letters are symbols for the names of the elements.

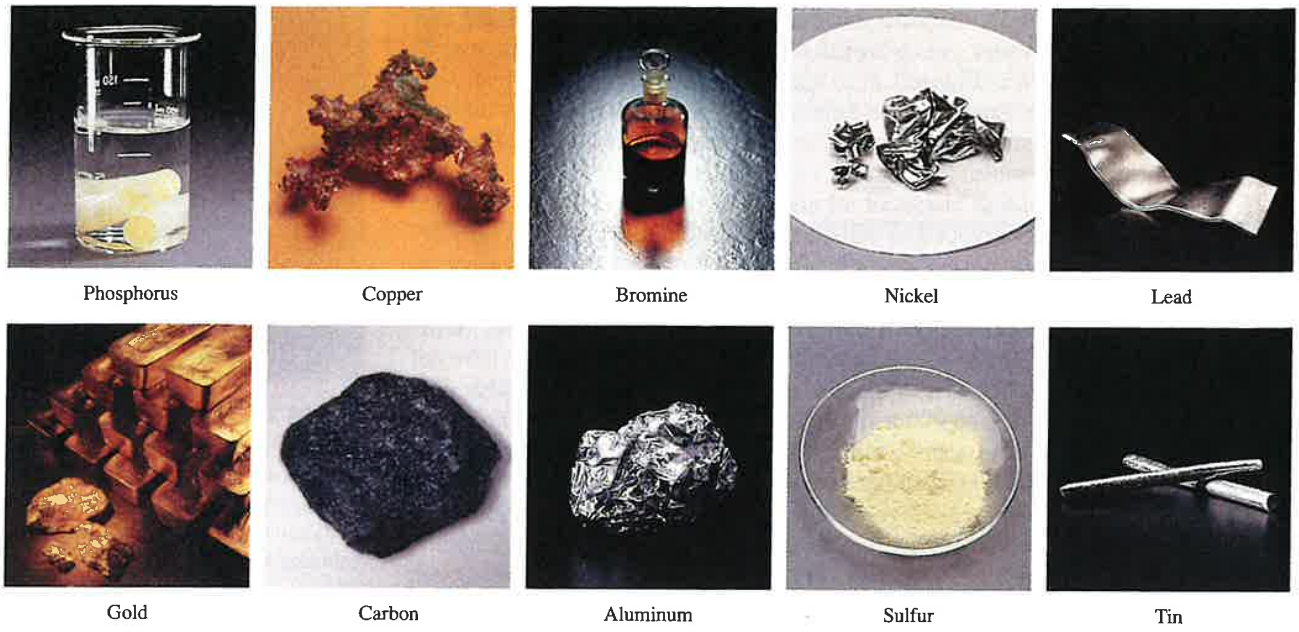
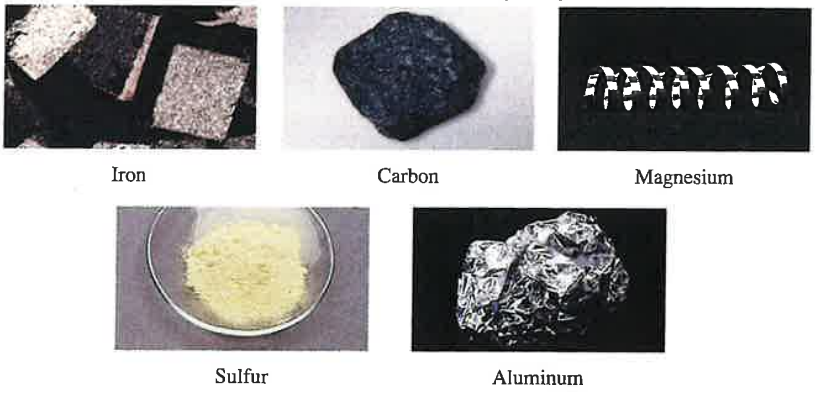


FIGURE 1.4 Some elements. Which of these are metals?

difference in appearance of the metals and nonmetals shown in Figure 1.4. Not all elements fit neatly into such categories. In Chapter 2 we'll discuss elements that have properties somewhere between metals and nonmetals.

EXAMPLE 1.1 | Metals and Nonmetals

Which of the elements pictured are metals? Why do you think so?



Solution:

Notice that three of the elements—iron, aluminum, and magnesium—have a luster; that is, they shine. They are metals. If you could handle and test the substances, you could use other properties, such as electrical conductivity, to distinguish between metals and nonmetals.

Practice Problem 1.1

Identify the nonmetals in Figure 1.4. Explain the characteristics you considered in making your decision.

Further Practice: Questions 1.29 and 1.30 at the end of the chapter

TABLE 1.1 | Symbols of Selected Elements

English Name	Original Name	Symbol	English Name	Original Name	Symbol
copper	cuprum	Cu	potassium	kalium	K
gold	aurum	Au	silver	argentum	Ag
iron	ferrum	Fe	sodium	natrium	Na
lead	plumbum	Pb	tin	stannum	Sn
mercury	hydrargyrum	Hg	tungsten	wolfram	W

To avoid having to write out the name of an element every time we refer to it, we use a system of symbols. An **element symbol** is a shorthand version of an element's longer name. Often, the symbol is one or two letters of the element's name (C for carbon, He for helium, Li for lithium). The first letter is uppercase, and the second letter, if present, is lowercase. When the names of two elements start with the same two first letters (magnesium and manganese, for example), the symbol uses the first letter and a later letter to distinguish them (Mg for magnesium, Mn for manganese).

For a few elements, the symbols are based on their Latin names or on names from other languages. These are listed in Table 1.1. Some recently synthesized elements have been named for famous scientists. Others have not been given permanent names. You'll find a list of the modern names and symbols on the inside front cover of this book.

To become familiar with the periodic table, you should learn the names and symbols for the first 36 elements, as well as the symbols for silver, tin, gold, mercury, and lead. Your instructor may ask you to learn others.

EXAMPLE 1.2 | Element Symbols

Potassium is a soft, silver-colored metal that reacts vigorously with water. Write the symbol for the element potassium.

Solution:

The symbol for potassium is K. In the periodic table, potassium is element 19 in group (column) IA (1) of the periodic table.

Practice Problem 1.2

(a) Lead is a soft, dull, silver-colored metal. Write the symbol for the element lead.
(b) The symbol for a common element used to make jewelry is Ag. What is the name of this element?

Further Practice: Questions 1.37 and 1.38 at the end of the chapter

Compounds A **compound,** sometimes called a *chemical compound,* is a substance composed of two or more elements combined in definite proportions. A compound has properties different from those of its component elements. For example, iron pyrite can be broken down into its component elements, iron and sulfur, but its characteristics are different from both (Figure 1.5). Anna and Bill saw many compounds that can be chemically separated into their component elements. Sand is a compound of silicon and oxygen. Water, as discussed earlier, is composed of hydrogen and oxygen. The cheese on their pizza contains many complex compounds, but each of the compounds contains carbon, hydrogen, oxygen, nitrogen, and a few other elements.



Iron pyrite



FIGURE 1.5 Iron pyrite is composed of the elements iron and sulfur. Iron is magnetic and can be separated from sulfur when the two exist as elements mixed together. Iron pyrite, a compound of iron and sulfur, is not magnetic.

Chemists represent compounds with formulas based on the symbols for the elements that are combined in the compound. (Chemical formulas are not the same as the mathematical formulas that may be familiar to you, such as $A = \pi r^2$ for the area of a circle.) A **chemical formula** describes the composition of a compound, using the symbols for the elements that make up the compound. Subscript numbers show the relative proportions of the elements in the compound. If no subscript number is given for an element in a formula, then you may assume that the element has a relative proportion of one. For example, water is known to consist of one unit of oxygen and two units of hydrogen. This compound is represented by the formula H_2O. Sodium chloride, the chemical compound commonly called table salt, contains equal portions of the elements sodium and chlorine. Its formula is therefore NaCl. We will discuss formulas in detail in Chapter 3.

Graphite leaves a mark similar to that made by dragging a rod of lead along a surface, so it was called lead. A hardness number indicates the relative amounts of graphite and clay in a pencil lead. A number 2 pencil is fairly soft, while a number 6 pencil is quite hard. Which has more graphite?

Mixtures Some forms of matter, such as pencil lead, do not have the same composition in every sample. (Pencil lead isn't the element lead. It is a *mixture* of graphite and clay.) A mixture consists of two or more elements or compounds. It is possible to separate mixtures into their component pure substances. The separation can be done physically, using procedures such as grinding, dissolving, or filtering. Chemical processes are not needed to separate mixtures.

We can illustrate the difference between pure substances and mixtures by looking at salt water. Water that has been purified is a pure substance that is composed of hydrogen and oxygen, always in the same proportions. Salt water, on the other hand, is water mixed with salt and many other substances in varying proportions. For example, the Great Salt Lake in Utah is approximately 10% salt, while the Dead Sea is about 30% salt. In either case, we can readily separate salt from water by evaporating the water (Figure 1.6).



FIGURE 1.6 To collect salt, water is diverted into large ponds. The water evaporates, leaving solid salt behind.

Mixtures differ in uniformity of composition. A **homogeneous mixture** has a uniform composition throughout and is often called a **solution.** Most solutions that we commonly encounter are composed of compounds dissolved in water. They are often clear. For example, a well-mixed sample of salt water prepared in a kitchen is uniform in appearance. The salt dissolved in it is invisible. Furthermore, any microscopically small portion of the sample would have the same composition as any other. The particles in the mixture might not be arranged in exactly the same pattern, but each sample, regardless of size, would have the same components in the same proportions.

A mixture that is not uniform throughout—a mixture of salt and pepper, for instance—is a **heterogeneous mixture.** Different samples have their components present in different proportions. Which of the things that Bill and Anna had for lunch is a homogeneous mixture? Which is heterogeneous? How about your own lunch? How can you tell?

We have considered a number of classes and subclasses of matter: mixtures, solutions, heterogeneous mixtures, pure substances, compounds, elements, metals, and nonmetals. A method for classifying matter into these categories is outlined in Figure 1.7. Note in the figure that yes or no answers to several questions distinguish one type of matter from another. First, we ask if the material can be separated physically. If so, then it is a mixture. If not, it must be a pure substance. If this substance can be decomposed (broken down into simpler substances) by chemical reactions, it is a compound. If it cannot, it is an element.

Not all solutions are liquids. For example, consider air that has been filtered to remove suspended solid particles. Filtered air is a gaseous solution containing a mixture of primarily oxygen and nitrogen gases, along with several other gases in lesser quantities. Solid solutions also exist and are called alloys. For example, the 14-carat gold used in rings is a solution of gold, silver, and copper.

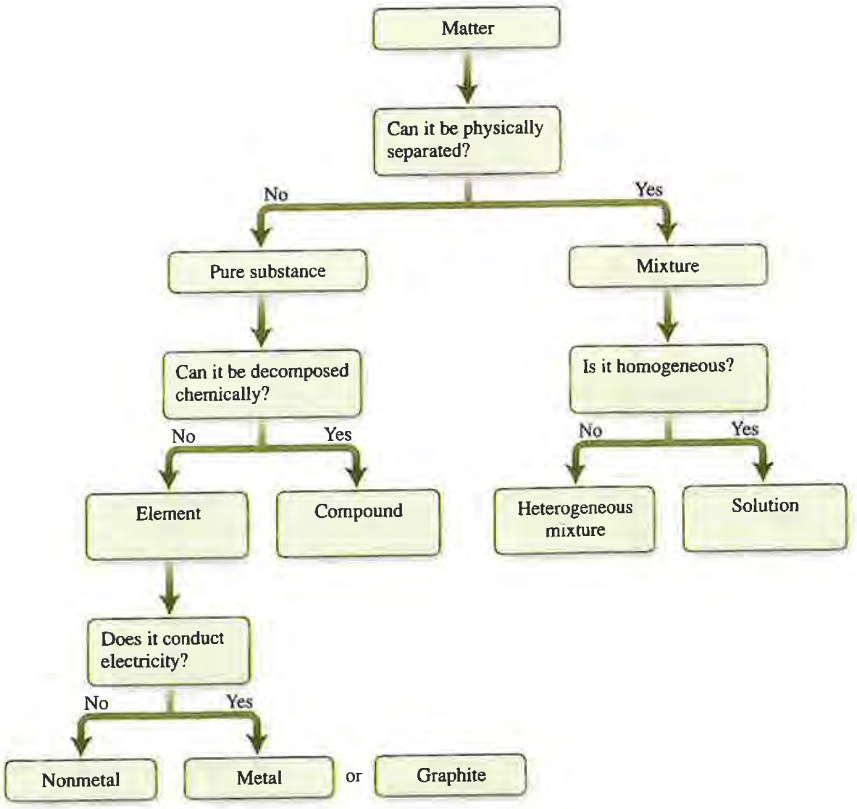


FIGURE 1.7 We can classify matter by answering the short series of questions in this flowchart.

EXAMPLE 1.3 | Elements, Compounds, and Mixtures

Which of the following pictures represent pure substances?











Solution:

The copper on the outside of the coin and the helium inside the balloons are pure substances. (However, the helium and balloons considered together provide an example of a mixture.)

Practice Problem 1.3

Which of the pictures represent mixtures? Which are heterogeneous? Which are homogeneous?

Further Practice: Questions 1.43 and 1.44 at the end of the chapter

Although chemists generally use color coding to distinguish between atoms of different elements in representations, the atoms themselves do not have colors. Macroscopic samples of matter may have color, but these colors do not usually match those used to represent atoms. In accurate representations, the sizes of the spheres change to reflect the relative differences in the sizes of atoms of different elements.

Representations of Matter

Chemists and other scientists view the world on several different levels. So far we have considered matter on a macroscopic scale. That is, we've discussed matter and phenomena we can see with our eyes. But simple observation is limited. Sometimes we cannot classify things merely by looking at them as Anna and Bill did. What do we do then? Chemists try to make sense of the structure of matter and its behavior on a scale that is much, much smaller than what we can see with our eyes.

Consider the copper pipe at the construction site, for example. If we could enlarge the tiniest unit that makes up the pipe, what would we see? Experimental evidence tells us copper is made up of discrete, spherical entities that all appear to be identical (Figure 1.8). Chemists identify these entities as atoms. An **atom** is the smallest unit of an element that has the chemical properties of that element. For example, we can imagine the helium inside a balloon as many, many atoms of helium, which we represent symbolically as He. In Figure 1.9, each sphere represents a single helium atom. Similarly, if we could magnify the structure of water, we would find two small hydrogen atoms bound separately to a single larger oxygen

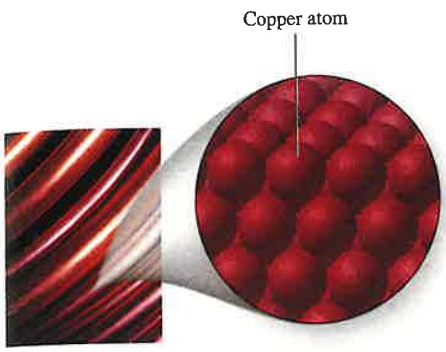


FIGURE 1.8 A copper pipe consists of a regular array of copper atoms.

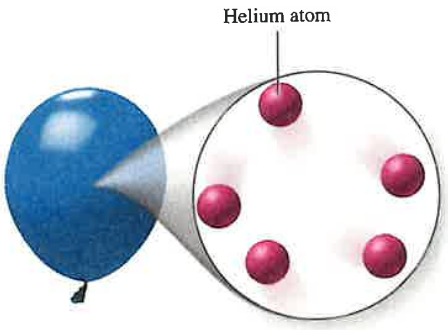


FIGURE 1.9 Helium atoms are present inside the balloon.

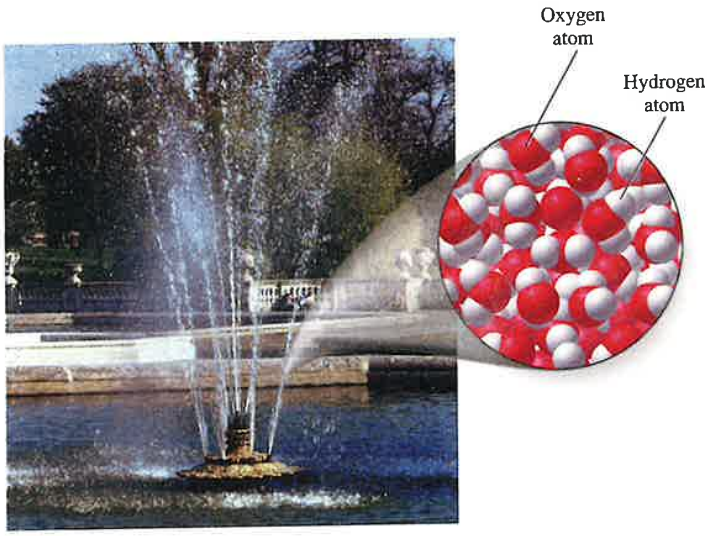


FIGURE 1.10 Molecules containing hydrogen atoms and oxygen atoms make up the water in the fountain.

atom. Such a combination of elemental units is a **molecule.** Molecules are made up of two or more atoms bound together in a discrete arrangement. Several molecules of water, H_2O, are shown in Figure 1.10, where the central red sphere represents an oxygen atom and the two smaller, white spheres stand for hydrogen atoms. (Some compounds do not exist as molecules. We will discuss them in Chapter 3.)

In addition to molecules of compounds, molecules can also be formed by the combination of atoms of only one element. For example, as shown in Figure 1.11, the oxygen we breathe consists of molecules of two oxygen atoms joined together. We represent oxygen molecules symbolically as O_2.

Chemists use many different ways to represent matter. Some are shown in Figure 1.12. Element symbols with subscripts represent a ratio of elements in a compound. One example is Figure 1.12B. To describe how the atoms are attached to one another, chemists often use lines and element symbols as shown in Figure 1.12C. In Figure 1.12D spheres represent the atoms, and sticks show how they are connected. Figure 1.12E represents how the atoms fit together and their relative sizes. Macroscopic, molecular-level, and symbolic representations like these all have their advantages, and sometimes one is more convenient than another. You'll use them all as you progress through this course.

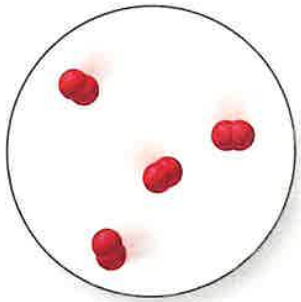


FIGURE 1.11 Oxygen molecules are made up of two interconnected oxygen atoms and are represented symbolically as O_2.

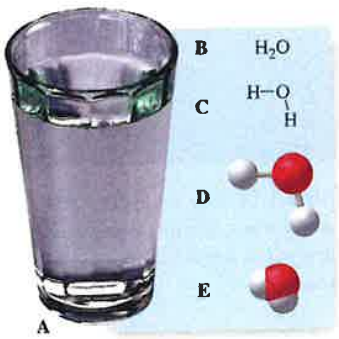
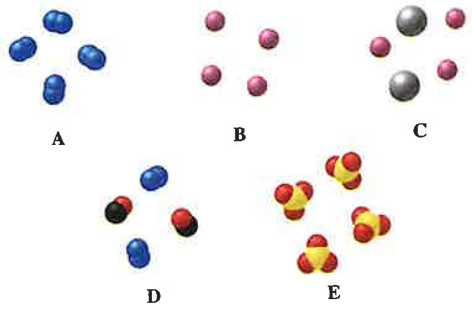


FIGURE 1.12 Different ways of representing water: (A) macroscopic, (B and C) symbolic, and (D and E) molecular.

EXAMPLE 1.4 | Representations of Matter

(a) Which of these images best represents a mixture of elements?



(b) If image A represents nitrogen, write its formula.

Solution:

(a) There are two mixtures represented in the images. Since the spheres (representing atoms) in image C have different colors and sizes, we can conclude that image C is a mixture of two elements. Image D is also a mixture, but it is a mixture of an element and a compound.
(b) The formula of the substance represented in image A is N_2. Note that two atoms are connected in the molecule.

Practice Problem 1.4

(a) Which of the images represents an element that exists as a molecule?
(b) If image E represents a compound of oxygen (red) and sulfur (yellow), what is its formula? (Write the symbol for sulfur first.)

Further Practice: Questions 1.49 and 1.50 at the end of the chapter

ANIMATION: Three States of Matter



FIGURE 1.13 Dry ice is the solid state of carbon dioxide. It converts from a gas to a solid at a very low temperature.

Some solids, called amorphous solids, do not have the high order that most crystalline solids have.

States of Matter

Earlier we considered the classification of matter based on composition. Let's look at a different way to classify matter: by its physical state. A **physical state** is a form that matter can take. The three most familiar to us are *solid, liquid,* and *gas.* Some substances, including some of those Anna and Bill observed, can be found in all three states under more or less ordinary conditions. Water, for example, can be a solid (ice), a liquid (flowing water), or a gas (water vapor) at environmental temperatures.

Other substances require extreme conditions to change from one state to another. For example, while carbon dioxide is a gas under normal conditions, it becomes a solid, called dry ice, at very low temperatures (Figure 1.13).

How do we know if a substance is in the solid, liquid, or gaseous state? Each state has characteristics that we can observe with our eyes and characteristics that are detectable or measurable at the molecular level. These characteristics are summarized in Table 1.2.

A **solid** has a fixed shape that is not related to the shape of the container holding it. When you place an iron pipe in a box, the pipe does not change shape. Some solids can be made to change shape if enough force is applied. However, if you try to squeeze a solid to make it smaller, you'll fail. A solid cannot be compressed because its particles are arranged in a tightly packed, highly ordered structure that does not include much free space into which they might be squeezed. Note the closely packed particles in the solid state of iron shown in Figure 1.14.

TABLE 1.2 | Characteristics of the Physical States of Matter

Solid	Liquid	Gas
fixed shape	shape of container (may or may not fill it)	shape of container (fills it)
its own volume	its own volume	volume of container
no volume change under pressure	slight volume change under pressure	large volume change under pressure
particles are fixed in place in a regular (crystalline) array	particles are randomly arranged and free to move about until they bump into one another	particles are widely separated and move independently of one another

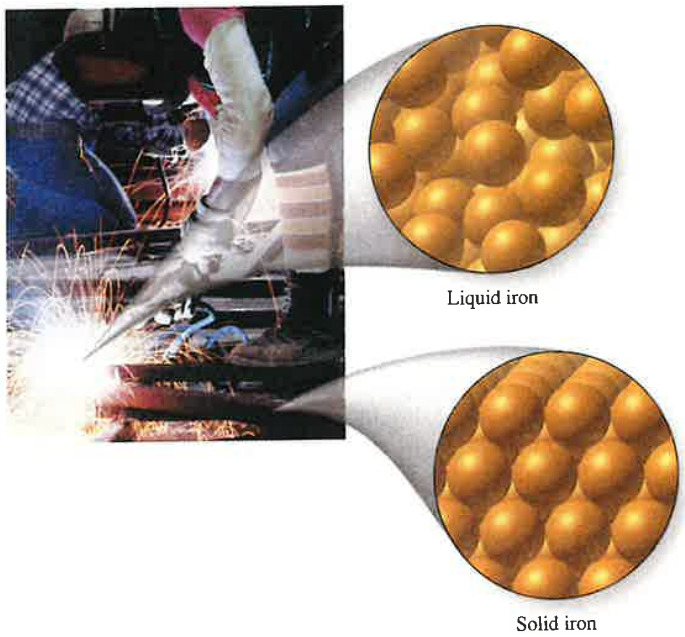


FIGURE 1.14 The liquid and solid states of iron.

A **liquid** is different from a solid in that it has no fixed shape. It takes the shape of the filled portion of its container, and it can be poured. Although they touch, the particles in a liquid are not arranged in ordered structures like those in a solid; they are free to move past one another. A liquid can be compressed slightly because its particles have a little free space between them. Note the differences between the liquid and solid states of iron shown in Figure 1.14.

A **gas** has no fixed shape; it adopts the shape of its container, expanding to fill the available space completely. A gas is easily compressed. When squeezed, gases can undergo large changes in volume. The particles of a gas are widely separated with much empty space between them. When a gas is compressed, the amount of space between the particles is reduced. This happens when pressure is applied, such as when a bicycle tire is filled with air, as shown in Figure 1.15. Another characteristic of gases is that they move through space quickly. When Bill and Anna smelled the pizza they had for lunch, they were detecting particles that migrated as gases from the source of the food to their noses. When gases cool sufficiently, they become liquids or even solids. This occurs, for example, when water vapor in the air liquefies on the surface of a cold glass. Note the differences between the liquid and gaseous states of water shown in Figure 1.16.

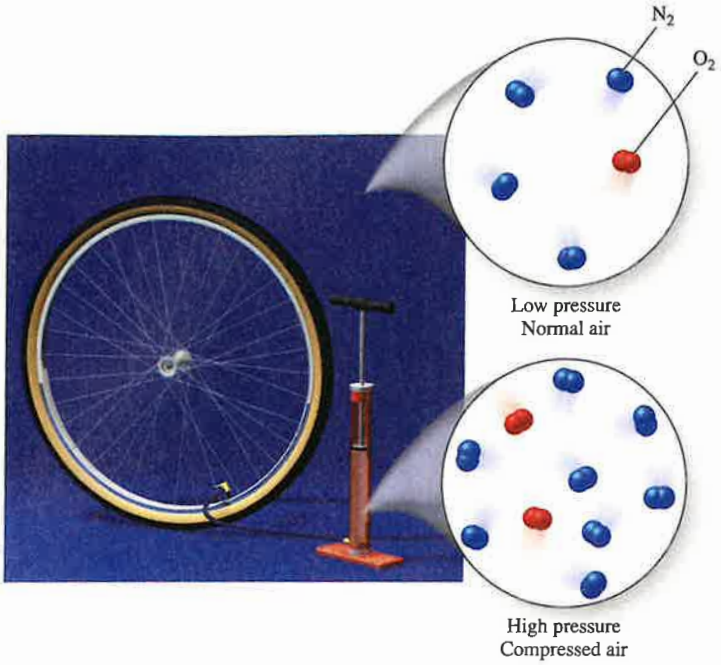


FIGURE 1.15 At the same temperature, a gas under high pressure has particles closer together than at low pressure. Notice that the composition (1 O_2:4 N_2) does not change with an increase in pressure.

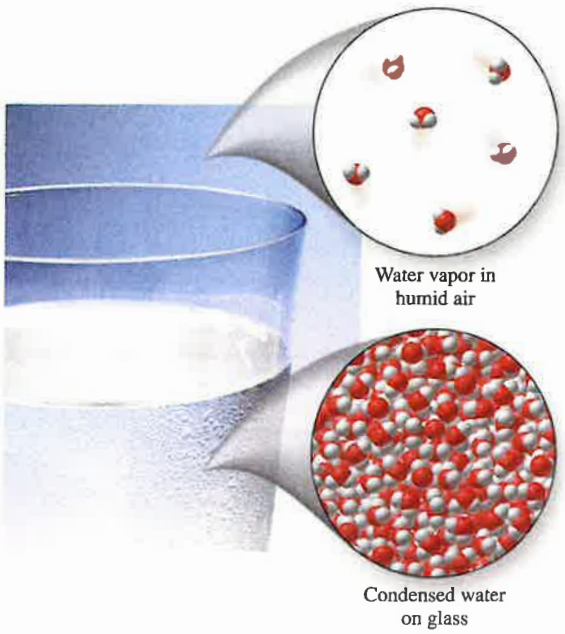


FIGURE 1.16 Water condenses from a gas to a liquid on a cold surface. Air molecules (e.g., oxygen and nitrogen) are not shown.

It is often convenient to show the physical state of a substance when representing it symbolically. For example, solid, liquid, and gaseous water can be represented as $H_2O(s)$, $H_2O(l)$, and $H_2O(g)$, respectively. The symbol (aq) represents an **aqueous solution**, a solution in which a substance is dissolved in water. A salt and water solution, for instance, can be written as $NaCl(aq)$. These symbols for physical state are listed in Table 1.3.

TABLE 1.3 | Symbols for Physical State

Physical State	Symbol	Example (bromine)
solid	(s)	$Br_2(s)$
liquid	(l)	$Br_2(l)$
gas	(g)	$Br_2(g)$
aqueous (dissolved in water)	(aq)	$Br_2(aq)$

1.2 | PHYSICAL AND CHEMICAL CHANGES AND PROPERTIES OF MATTER

Bill and Anna observed some of the *properties* of matter, including *changes* in matter. Their observations could be either *qualitative*, based on some quality of the matter; or *quantitative*, based on a numerical value. When making qualitative observations, they described color, shape, texture, shininess, and physical state. Quantitative observations are different. They are numbers or measurements, and they must be carefully made and carefully reported.

Since quantitative data used to describe matter can involve both very large and very small numbers, it is often useful to express such numbers in *scientific* or *exponential notation.* Math Toolbox 1.1 (located at the end of this chapter) provides a review of this notation. In addition, it is necessary to express numbers in such a way as to indicate how accurately the value is known and how precisely it has been measured. The use of *significant figures* to properly express numerical values is presented in Math Toolbox 1.2.

MATH TOOLBOX 1.1

MATH TOOLBOX 1.2

Physical Properties

When reporting qualitative data, we can classify properties as either physical or chemical. When Bill and Anna observed the color, shape, texture, shininess, and physical state of things around them, they were noting their physical properties. A **physical property** is a characteristic that we can observe or measure without changing the composition of a substance. Other examples of physical properties are odor, taste, hardness, mass, volume, density, magnetism, conductivity, and the temperatures at which a substance changes from one physical state to another. Let's take a close look at mass, volume, density, and temperature. These four properties are *quantitative;* they involve numerical values.

Mass Recall that mass is a measure of the quantity of matter. We usually measure the mass of an object by weighing it on a balance. In chemistry, masses are often reported in units of grams (g). Large masses, like people or elephants, may be reported in units of kilograms (kg); and small masses, such as salt crystals or impurities in water, may be reported in units of milligrams (mg) or micrograms (μg), as shown in Figure 1.17. (Math Toolbox 1.3 summarizes the relationships among units such as these.) Sometimes the mass of something is reported in grams, but we might want to know the mass in another mass unit such as milligrams or kilograms. We can easily convert a measurement from one unit to another if we know the relationship between the units. Tables 1.4 and 1.5 summarize common relationships between metric and English units. Example 1.5 shows how to convert between mass units. (See Math Toolbox 1.3 for more information on unit conversions.)

MATH TOOLBOX 1.3



Mass: 50 mg, 0.05 g, or 5×10^{-5} kg



Mass: 7×10^{7} mg, 7×10^{4} g, or 70 kg

FIGURE 1.17 A salt crystal has a mass of about 50 mg, while a person has a mass of about 70 kg.

TABLE 1.4 | Metric Conversions

Prefix	Factor	Symbol
giga	10^{9}	G
mega	10^{6}	M
kilo	10^{3}	k
deci	10^{-1}	d
centi	10^{-2}	c
milli	10^{-3}	m
micro	10^{-6}	μ
nano	10^{-9}	n
pico	10^{-12}	p

TABLE 1.5 | Some English-Metric Conversions

English Unit	Metric Unit
1 lb = 16 oz	453.6 g
1 in	2.54 cm (exactly)
1 yd	0.9144 m
1 mi	1.609 km
1 fluid oz	29.57 mL
1 qt	0.9464 L
1 gal	3.785 L
1 ft^3	28.32 L

EXAMPLE 1.5 | Units of Mass

Anna and Bill notice that there are 50.0 mg of sodium in the soda they bought to go with their lunch. How many grams of sodium are present in the can of soda? How many pounds?

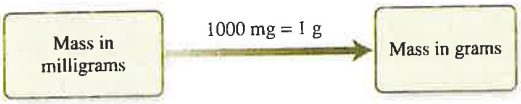
Solution:

MATH TOOLBOX 1.3

One way to solve this problem uses the dimensional-analysis approach. Consult Math Toolbox 1.3 for details. The general approach to solving the first part of the problem can be summarized by the following diagram:



The mass in milligrams has to be converted to the mass in grams. We need to find a relationship between these two quantities: 10^{-3} g = 1 mg or 1 g = 1000 mg (obtained from Table 1.4). Using this relationship, we get the following conversion:



We use the equivalence to set up possible conversion ratios:

$$\frac{1\text{ g}}{1000\text{ mg}} \quad \text{and} \quad \frac{1000\text{ mg}}{1\text{ g}}$$

To convert milligrams to grams, we can multiply 50.0 mg by the ratio (conversion factor) that will allow like units to cancel:

$$\text{Mass in mg} = 50.0\,\cancel{\text{mg}} \times \frac{1\text{ g}}{1000\,\cancel{\text{mg}}} = 0.0500\text{ g}$$

MATH TOOLBOX 1.2

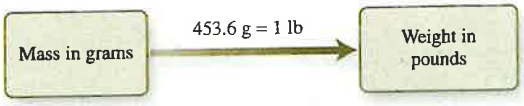
MATH TOOLBOX 1.1

Note that the milligram units cancel to leave the appropriate unit of grams. Also notice that the answer is reported to three significant figures because the measured quantity (50.0 mg) is reported to three significant figures and the other numbers in the calculation (1 g and 1000 mg) are exact quantities. (Consult Math Toolbox 1.2 for details about significant figures.) For convenience, we could report the answer in scientific notation: 5.00×10^{-2}. (See Math Toolbox 1.1 for a discussion of scientific notation.)

The second part of the question asks you to convert milligrams to pounds:



There isn't a direct relationship between milligrams and pounds listed in Tables 1.4 and 1.5. However, Table 1.5 lists a relationship between pounds and grams: 1 lb = 453.6 g. We can convert the grams we found in the first part of this example to pounds using the relationship summarized in the following diagram:



The ratios for converting between grams and pounds are

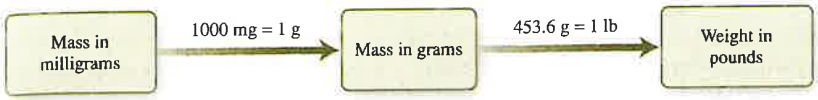
$$\frac{1\text{ lb}}{453.6\text{ g}} \quad \text{and} \quad \frac{453.6\text{ g}}{1\text{ lb}}$$

To convert grams to pounds, we can multiply 0.0500 g by the ratio (conversion factor) that will allow like units to cancel:

$$\text{Weight in pounds} = 0.0500\,\cancel{g} \times \frac{1\text{ lb}}{453.6\,\cancel{g}} = 1.10 \times 10^{-4}\text{ lb}$$

Does this answer make sense? Yes, it does. There are a lot of grams (453.6) in a pound, so we would expect the answer to be very small. Would the answer 22.7 lb make sense? No.

Without a single conversion from milligrams to pounds, the problem we just solved involves multiple steps:



The sequence of steps can be summarized as:

$$\text{Weight in pounds} = 50.0\,\cancel{mg} \times \frac{1\,\cancel{g}}{1000\,\cancel{mg}} \times \frac{1\text{ lb}}{453.6\,\cancel{g}} = 1.10 \times 10^{-4}\text{ lb}$$

Practice Problem 1.5

Anna and Bill see an aluminum recycling truck pass by on their way to class. If there are 765 lb of aluminum in the truck how many grams are there? How many kilograms?

Further Practice: Questions 1.65 and 1.66 at the end of the chapter

Volume **Volume** is the amount of space a substance occupies. We can determine the volume of a cube by measuring its length, width, and height and then multiplying them. For example, the volume of a cube that is 2.0 centimeters (cm) on each side is 8.0 cubic centimeters (cm^3):

$$\text{Volume of a cube} = \text{length} \times \text{width} \times \text{height}$$

$$\text{Volume} = 2.0\text{ cm} \times 2.0\text{ cm} \times 2.0\text{ cm} = 8.0\text{ cm}^3$$

The volumes of liquids are usually measured in units of liters (L) or milliliters (mL), as shown in Figure 1.18. One cubic centimeter is equal to 1 mL, so the volume of 8.0 cm^3 could also be reported as 8.0 mL. Larger volumes, such as big bottles of soda, are usually reported in liters. A 1-L bottle of soda contains 1000 mL. Example 1.6 shows how to convert between volume units.

If you need to determine the volume of a sphere, the relationship between volume and radius is $V = \frac{4}{3}\pi r^3$.



FIGURE 1.18 Some 250-mL, 500-mL, and 1-L containers.

EXAMPLE 1.6 | Units of Volume

For lunch, Anna and Bill had 12-ounce (oz) cans of soda. What is the volume of a 12.0-oz can of soda in units of milliliters? What is its volume in units of liters?

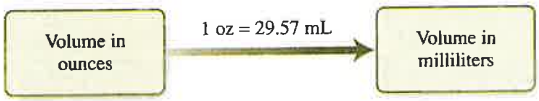
Solution:

MATH TOOLBOX 1.3

To solve this problem using the dimensional-analysis approach (see Math Toolbox 1.3), we determine if there is a relationship between fluid ounces and milliliters:



To convert fluid ounces to milliliters we use the following relationship from Table 1.5: 1 oz = 29.57 mL.



We use the equivalence to set up possible conversion ratios:

$$\frac{29.57\ \text{mL}}{1\ \text{oz}} \quad \text{and} \quad \frac{1\ \text{oz}}{29.57\ \text{mL}}$$

To convert ounces to milliliters, we can multiply 12.0 oz by the ratio (conversion factor) that will allow like units to cancel:

$$\text{Volume in milliliters} = 12.0\ \cancel{\text{oz}} \times \frac{29.57\ \text{mL}}{1\ \cancel{\text{oz}}} = 355\ \text{mL}$$

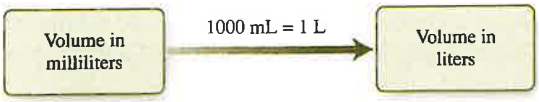
MATH TOOLBOX 1.2

The answer is reported to three significant figures, because the quantity we're given (12.0 oz) has three significant figures. Consult Math Toolbox 1.2 for details.

The second part of this problem asks you to convert milliliters to liters:



To convert volume in milliliters to volume in liters, we use the following relationship from Table 1.4: 1 mL = 10^{-3} L or 1000 mL = 1 L.



This is a unit conversion similar to the conversion we just did for mass. The ratios for converting between milliliters and liters are

$$\frac{1\ \text{L}}{1000\ \text{mL}} \quad \text{and} \quad \frac{1000\ \text{mL}}{1\ \text{L}}$$

To convert from milliliters to liters, we can multiply 355 mL by the conversion factor that allows like units to cancel:

$$\text{Volume in L} = 355\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.355\ \text{L}$$

Without a single conversion from ounces to liters, the problem we just solved involves multiple steps:



The sequence of steps can be summarized as:

$$\text{Volume in liters} = 12.0 \cancel{\text{oz}} \times \frac{29.57 \cancel{\text{mL}}}{1 \cancel{\text{oz}}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.355 \text{ L}$$

Practice Problem 1.6

Anna and Bill saw some balloons outside the bookstore. The volume of gas inside one of the helium balloons was 4.60 L. What is the volume of gas in units of milliliters? In units of cubic centimeters? In units of gallons (4 qt = 1 gal)?

Further Practice: Questions 1.69 and 1.70 at the end of the chapter

Density The **density** of an object is the ratio of its mass to its volume. While mass and volume both depend on the size of the object or sample, density does not. Density is an unvarying property of a substance no matter how much of it is present, as long as temperature and pressure are constant. The densities of a few substances are listed in Table 1.6.

As Anna and Bill noted when they observed the fountain, a copper coin sinks in water. It sinks because copper (and the other metals in a penny) have a greater density than water. Conversely, air bubbles, just like other gases, rise to the top of water because gases are less dense than liquids. Oil floats on water for this same reason.

The density column in Figure 1.19 shows a variety of liquids with different densities. Which liquid has the greatest density? Which is the least dense?

If we compare *equal volumes* of two different substances, such as aluminum and gold, as shown in Figure 1.20, the substance with the greater mass has the greater density. How, though, can we compare densities if we do not have equal volumes? The mathematical relationship of mass, volume, and density reveals the answer:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

For example, a 1.0-cm^3 sample of copper has a mass of 8.9 g. An 8.0-cm^3 sample of copper has a mass of 71 g. A 27-cm^3 sample of copper has a mass of 240 g. In all these samples (Figure 1.21), the mass of copper divided by its volume is 8.9 g/cm^3. This is the density of copper. If we know the mass and volume of an object, we can determine its density by substituting directly into the density equation.



FIGURE 1.19 The densities of antifreeze, corn oil, dish detergent, maple syrup, shampoo, and water in g/mL are 1.13, 0.93, 1.03, 1.32, 1.01, and 1.00, respectively, Which layer is which substance?



FIGURE 1.20 Gold (Au) has a greater density than aluminum (Al) because gold has a greater mass per unit volume.



FIGURE 1.21 The density of copper is 8.9 g/cm^3. All three samples have the same ratio of mass to volume.

TABLE 1.6 | Densities of Some Common Substances

Substance	Physical State	Density (g/mL)*
helium	gas	0.000178
oxygen	gas	0.00143
cooking oil	liquid	0.92
water	liquid	1.00
mercury	liquid	13.6
gold	solid	19.3
copper	solid	8.92
zinc	solid	7.14
ice	solid	0.92

*At room temperature and at normal atmospheric pressure, except gases at 0 degrees Celsius (°C) and water at 4°C.

ANIMATION: Density of Liquids and Solids

Additionally, if we know the density of a substance and its mass in our sample, we can determine its volume. For example, suppose we want to know the volume occupied by 100 g of copper. Should the volume be greater than or less than 100 cm^3? There are many approaches to this problem. One way is to rearrange the density equation to solve for volume. Another way is to solve for the unknown volume in a set of equivalent ratios because density is a ratio of mass and volume that is constant for a given substance at a particular temperature. Both of these methods are shown in Example 1.7.

These samples of metals have the same mass. Which has the greater density?



MATH TOOLBOX 1.3

A can of diet cola floats in water, but a can of regular cola sinks. Suggest a reason why. How can you use this information to quickly select your preferred type of soft drink from a cooler filled with ice water at a party?



EXAMPLE 1.7 | Density, Volume, and Mass

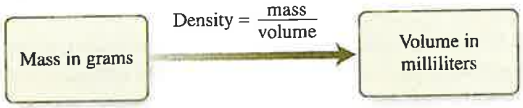
What is the volume of 100.0 g of copper? The density of copper is 8.9 g/cm^3.

Solution:

We need to carry out the following conversion:



The relationship between mass and volume is given by density:



First, we rearrange the density equation to get volume on one side by itself. This manipulation involves cross multiplication, which is described in Math Toolbox 1.3 (*Ratio Approach*). In the expression for density there is an implied 1:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$\frac{\text{Density}}{1} = \frac{\text{mass}}{\text{volume}}$$

Cross multiplying this density expression, we get:

$$\text{Density} \times \text{volume} = \text{mass} \times 1$$

Since we are trying to find the volume, we want to isolate it on one side of the equation. We can do this by dividing both sides by the density. (We'll also drop the "× 1" because any quantity times 1 is that quantity.)

$$\frac{\cancel{\text{Density}} \times \text{volume}}{\cancel{\text{Density}}} = \frac{\text{mass}}{\text{density}}$$

Now we have an expression that solves for the volume:

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

Then, we substitute the known values of mass and density into the equation and solve for the value of volume:

$$\text{Volume} = \frac{100.0\ \text{g}}{8.9\ \text{g/cm}^3} = 11\ \text{cm}^3$$

In a second approach to this problem, consider that since the density of copper is always the same, the ratio of mass to volume is the same for both what we know and what we don't:

$$\frac{8.9\ \text{g}}{1\ \text{cm}^3} = \frac{100.0\ \text{g}}{x\ \text{cm}^3}$$

Cross multiply to solve for x:

$$x\ \text{cm}^3 = \frac{(1\ \text{cm}^3) \times (100.0\ \cancel{\text{g}})}{8.9\ \cancel{\text{g}}} = 11\ \text{cm}^3$$

In both approaches, the gram units cancel to give the expected volume unit of cm^3.

There is yet another approach to solving this problem that involves using density as a conversion factor:

$$\text{Volume} = 100.0\ \cancel{\text{g}} \times \frac{1\ \text{cm}^3}{8.9\ \cancel{\text{g}}} = 11\ \text{cm}^3$$

Does the answer make sense? Yes. The density tells us that 8.9 g of copper occupy a volume of 1 cm^3. The mass given, 100.0 g, is over 10 times greater than 8.9 so we would expect it to occupy a volume that is over 10 times greater than 1 cm^3.

Practice Problem 1.7

Solve the following problems.

(a) The density of pure gold is 19.3 g/cm^3. What is the volume of 1.00 g of pure gold?

(b) 14-Carat gold is a homogeneous mixture of metals containing 58% gold by mass. The other 42% is a mixture of silver and copper. Silver and copper are both less dense than gold. Which of the following could be the mass of 1.00 cm^3 of 14-carat gold: 16.0 g, 19.3 g, or 23.0 g?

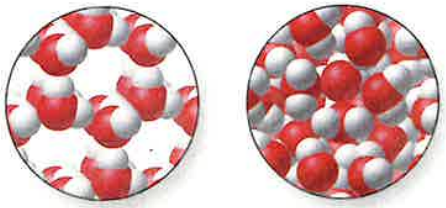
Further Practice: Questions 1.75 and 1.76 at the end of the chapter

Why do substances have different densities? Gases, in general, have very low densities because gas particles spread out and occupy large volumes. Metals tend to have high densities because their atoms pack together efficiently. Because ice floats on water, we can infer that water in its solid form must have a lesser density than water in its liquid form. Example 1.8 shows how to use molecular pictures to predict relative densities.

Water is unique among liquids because its solid form (ice) floats on its liquid form. This results from the relatively open structure adopted by water molecules in the solid state. What would happen to fish during the winter if ice were like other solids that sink in their liquid form?

EXAMPLE 1.8 | Explanations for Density

How do the molecular diagrams of ice and water help explain why ice is less dense than water?



Ice Liquid water

ANIMATION: Unique Properties of Water

Solution:

In ice, the H_2O molecules have more space between them than in liquid water. The total volume occupied by a given number of molecules is greater in ice. Because density is a ratio of mass to volume, the larger volume accounts for the lower density.

Practice Problem 1.8

Helium balloons rise in air, which is a mixture of oxygen and nitrogen molecules, so we know helium is less dense than air. Look at the molecular-level diagrams of helium and carbon dioxide. Predict whether a helium balloon rises or falls in an atmosphere of carbon dioxide.



Helium Carbon dioxide

Further Practice: Questions 1.79 and 1.80 at the end of the chapter

Temperature Bill and Anna weren't happy with their lunches. The pizza was cold and their sodas were warm. When we make such comparisons, we are observing relative temperatures. **Temperature** is a measure of how hot or cold something is relative to some standard. We measure temperature with a thermometer.

In the United States, we often use the Fahrenheit scale to measure body temperature and air temperature. Fahrenheit is rarely used in science. Two other temperature scales are standard: the Celsius scale and the kelvin scale. The relationships between the three temperature scales, Fahrenheit (°F), Celsius (°C), and kelvin (K), are shown in Figure 1.22.

Temperatures are written differently for the different scales. While Celsius and Fahrenheit use the superscript ° to indicate degrees, the kelvin scale does not. The unit is written as K (the capital letter), but temperatures are measured in kelvins (lowercase).

Another property of matter that is independent of sample size is the temperature at which the substance changes from one physical state to another. The *boiling point* is the temperature at which the liquid form of a substance changes to the gaseous form. At the *melting point,* the substance changes from a solid to a liquid. Between these two temperatures, the substance is normally in its liquid state. For example, on the Celsius scale, the boiling point of water is 100°C. Water

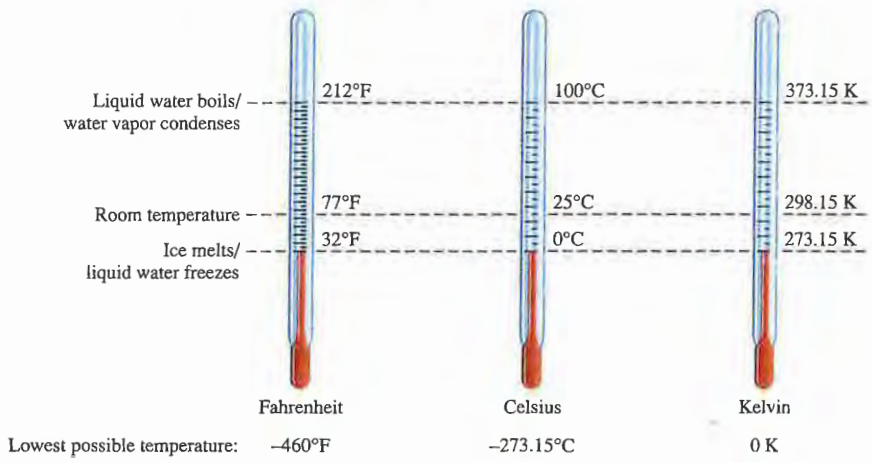


FIGURE 1.22 The Fahrenheit, Celsius, and kelvin temperature scales.

melts (or freezes, depending on its original state) at 0°C. On the kelvin scale, these values are 373.15 K and 273.15 K, respectively. On the Fahrenheit scale, they are 212°F and 32°F, respectively.

There are no negative values on the kelvin scale. It is an *absolute temperature scale* because its zero point is the lowest possible temperature observable in the universe. This value is absolute zero, which is equivalent to –273.15°C. The temperature increments on the kelvin scale are the same as those on the Celsius scale. The *difference* in temperature between the boiling point of water and the freezing point of water is 100 in both the Celsius (100°C – 0°C) and kelvin (373.15 K – 273.15 K) scales, while the difference is 180 on the Fahrenheit scale (212°F – 32°F). Because the temperature in kelvins is always 273.15 greater than the temperature in degrees Celsius, we can easily convert between them:

$$T_K = T_{°C} + 273.15$$

When converting between the Fahrenheit and Celsius scales, the calculation is more complicated, because the degree increments are not equal:

$$T_{°F} = 1.8(T_{°C}) + 32$$

The equation can be rearranged, solving for degrees Celsius:

$$T_{°C} = \frac{T_{°F} - 32}{1.8}$$

EXAMPLE 1.9 | Units of Temperature

The melting point of copper is 1083°C. Above what temperature, in kelvins and degrees Fahrenheit, is copper a liquid?

Solution:

Copper becomes a liquid above its melting point. In units of kelvin this temperature is

$$T_K = T_{°C} + 273.15$$

We substitute the value of the Celsius temperature into the expression and solve for the temperature in kelvins:

$$T_K = 1083 + 273.15 = 1356 \text{ K}$$

To convert degrees Celsius to degrees Fahrenheit we use the equation:

$$T_{°F} = 1.8(T_{°C}) + 32$$

Substituting the temperature in degrees Celsius we get:

$$T_{°F} = 1.8(1083) + 32 = 1981°\text{F}$$

Practice Problem 1.9

(a) The boiling point of acetylene is –28.1°C. Below what temperature, in kelvins and degrees Fahrenheit, is acetylene a liquid?
(b) The boiling point of helium is 4 K. Below what temperature, in degrees Celsius, is helium a liquid?
(c) Human body temperature is normally 98.6°F. What is this temperature on the Celsius and kelvin scales?

Further Practice: Questions 1.83 and 1.84 at the end of the chapter

Physical Changes

A process that changes the physical properties of a substance *without changing its chemical composition* is a **physical change.** For example, we can change liquid water to water vapor by heating it. This change from a liquid to a gas, called *boiling* or *vaporization,* is a physical change, since both forms involve the same chemical substance, water (H_2O).

To represent such changes, we can refine the symbolic representations we developed for elements and for compounds. We write the chemical formula for the initial condition and composition of the matter we are considering, then an arrow, and finally the chemical formula for the final condition and composition. The arrow is used to show that a change has occurred and in which direction. Using this symbolism, the change of water from a liquid to a gas would be represented as

$$H_2O(l) \longrightarrow H_2O(g)$$

The molecular and symbolic representations in Figure 1.23 show that the water molecules do not themselves change, but their physical state does. All the processes that change water from one physical state into another are summarized in Figure 1.24.

Another example of a physical change is the separation of different substances in a mixture. For example, a magnet divides magnetic materials from nonmagnetic materials without changing their identities. A filter separates solid materials from liquid substances without changing either one chemically.

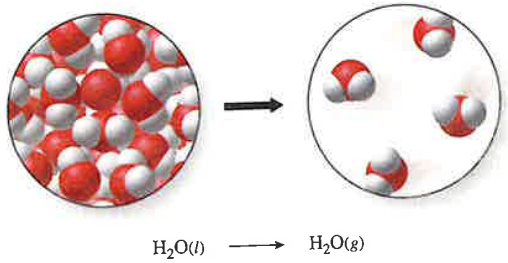


FIGURE 1.23 Molecular-level and symbolic representations of the evaporation of water.

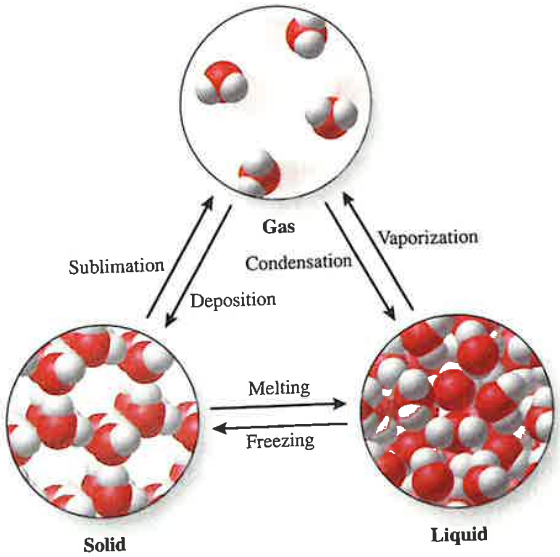
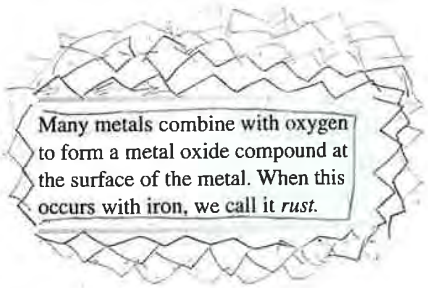


FIGURE 1.24 The physical states of solid, liquid, and gas can all change into one another either directly or by going through two changes of state. The names of these processes are shown here next to arrows that designate the direction of the change.

Chemical Changes

Remember the pennies in the fountain that Anna and Bill observed? Some were shiny and others looked dingy and brown. They might describe these less-shiny pennies as "tarnished." The pennies have undergone a **chemical change**, a process in which one or more substances are converted into one or more new substances. When pennies tarnish, some of the copper and zinc metal atoms in them combine with oxygen, forming compounds called metal oxides. The compounds are chemically different from either of the elements that formed them.


Many metals combine with oxygen to form a metal oxide compound at the surface of the metal. When this occurs with iron, we call it *rust*.

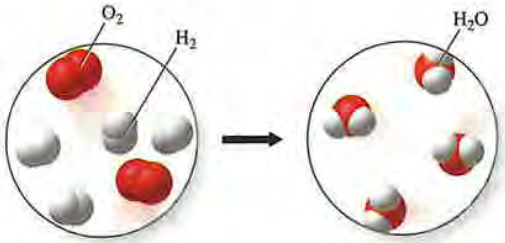
Suppose we clean a tarnished penny. Is the process a physical or a chemical change? It can be either. If you simply rub off the metal oxide coating with an eraser, the change is physical. Most penny collectors, however, prefer a chemical change that removes less metal. Rubbing ketchup on a penny is a great way to make it shiny. The vinegar in the ketchup reacts chemically with the metal oxides, freeing them from the surface of the penny. When the penny is rinsed, the result of the chemical change is easy to see.

Anna and Bill observed other examples of chemical change during their campus walk. When gasoline-powered cars burn fuel, a chemical change occurs. The gasoline reacts with oxygen to form carbon dioxide and water vapor. This chemical change releases the energy that runs the car. Chemical changes that involve burning are often accompanied by the release of energy. Anna and Bill also observed vehicles that run on alternative fuels. In hydrogen-powered vehicles, the hydrogen fuel combines with oxygen to form water vapor—and to release a lot of energy. The molecular-level and symbolic representations for this chemical change are shown in Figure 1.25. A chemical change is often called a **chemical reaction**. What are some examples of chemical reactions that you can observe around you?

Chemical Properties

The copper and zinc in a penny, the gasoline in a car, and the hydrogen in an alternative-fuel vehicle all share a common chemical property: They react with oxygen. However, they differ in how they react and what products they form. Only the latter two release sufficient energy rapidly enough to make their use as fuels possible.

A **chemical property** of a substance is defined by what it is composed of and what chemical changes it can undergo. For example, let's compare hydrogen and helium. Although they have similar physical properties (colorless gases, similar densities), their chemical properties are very different. While hydrogen reacts with many other elements and compounds, helium is considered *inert* (Figure 1.26). It has not yet been shown to react with any other element or compound.



$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

FIGURE 1.25 A chemical change occurs when the atoms in H_2 and O_2 rearrange to form H_2O.



FIGURE 1.26 The Hindenburg was a giant, rigid balloon filled with hydrogen gas. In 1937, it was destroyed when its hydrogen caught fire. Today, blimps are filled with helium, an inert gas that will not explode.

EXAMPLE 1.10 | Physical and Chemical Changes

Which of the following are physical changes and which are chemical changes?

(a) evaporation
(b) burning methane gas to form carbon dioxide and water
(c) using a magnet to separate metal and plastic paperclips
(d) rusting (the conversion of iron to iron oxide)

Solution:

(a) Evaporation is a physical change because it involves only a change of state.
(b) Burning methane gas is a chemical change because new substances form.
(c) Separating components of a mixture is a physical change.
(d) Rusting is a chemical change because a new substance forms.

Practice Problem 1.10

Which of the following are physical properties and which are chemical properties?

(a) boiling point of ethanol
(b) ability of propane to burn
(c) tendency for silver to tarnish
(d) density of aluminum

Further Practice: Questions 1.89 and 1.90 at the end of the chapter

Sometimes simple observation cannot tell us whether a change is chemical or physical. For example, bubbles appear when baking soda and vinegar mix. Bubbles also appear when water boils, but the change that produces the bubbles is different in these two cases. Baking soda and vinegar release bubbles because a chemical change takes place. They react to form carbon dioxide gas. However, when we warm water in a pan on the stove, small bubbles rise due to the release of dissolved air

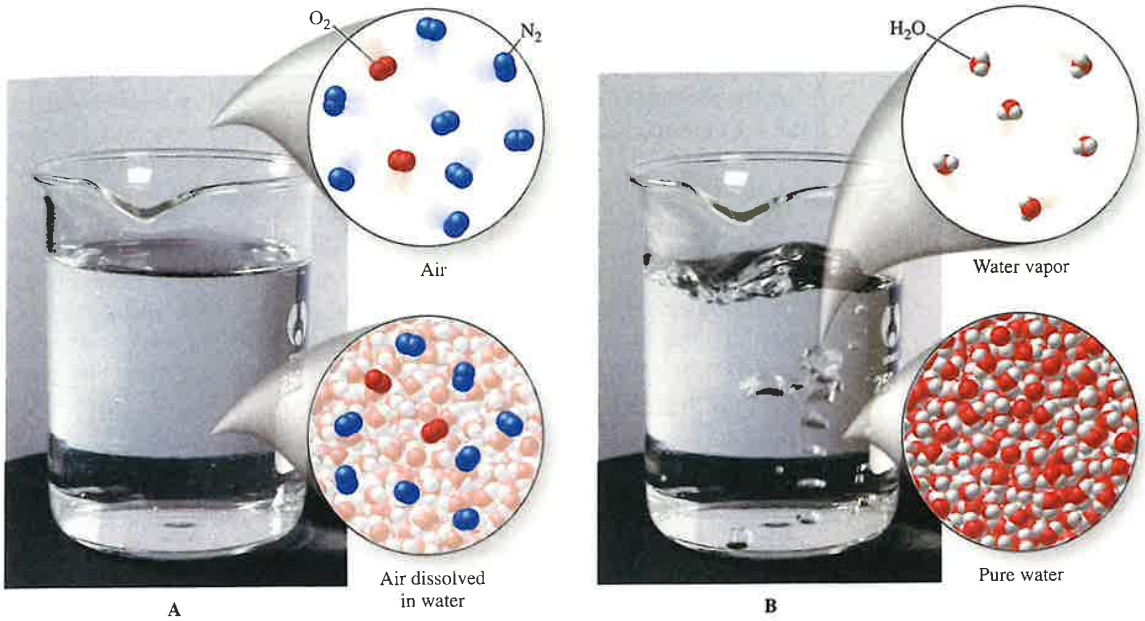


FIGURE 1.27 Water contains small amounts of dissolved nitrogen and oxygen gases. (A) When heated, these molecules go to the gaseous state in bubbles that rise to the surface. (B) When the water begins to boil, it no longer contains dissolved gases, and the bubbles contain gaseous water.

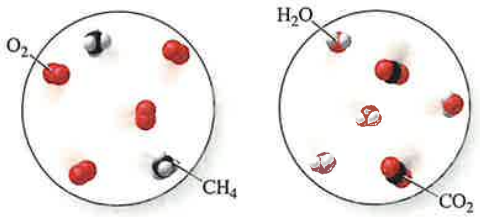
(mostly oxygen and nitrogen gas) from the water (before the water starts to boil). This process is only a physical change. If we could look at the nitrogen and oxygen molecules, as shown in Figure 1.27A, we would see that they are the same whether they are dissolved in water or not. When these molecules are dissolved in water, they form a homogeneous mixture with it. Heating the water merely separates the oxygen and nitrogen molecules from the water molecules. If we continue to heat the water to boiling, larger bubbles form and then rise from the bottom, as shown in Figure 1.27B. These bubbles are gaseous water, or water vapor. The result of the physical change can be represented symbolically as

$$H_2O(l) \longrightarrow H_2O(g)$$

Water molecules surround solute particles in aqueous solutions. For clarity in seeing the particles, this book shows water molecules faded in the background as in Figure 1.27A.

EXAMPLE 1.11 | Physical and Chemical Changes

Do the following molecular-level images represent a chemical change or a physical change?

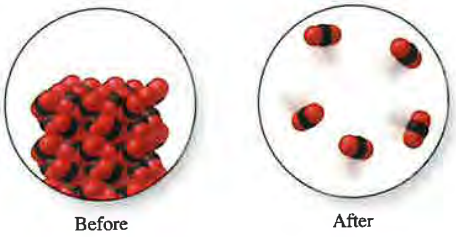


Solution:

The substances after the change have a different composition than the substances before the change. Therefore, this is a chemical change.

Practice Problem 1.11

Do the following molecular-level images represent a chemical change or a physical change?



Further Practice: Questions 1.93 and 1.94 at the end of the chapter

1.3 | ENERGY AND ENERGY CHANGES

Physical and chemical changes involve energy. Energy is hard to define, but we see and feel evidence of it when something moves or changes temperature. At the construction site, Anna and Bill saw a worker pushing a wheelbarrow up a ramp. If released at the top of the ramp, the wheelbarrow would roll back down, converting energy from one form to another in the process. This release of energy is related to the spontaneous process of rolling down the ramp. (A spontaneous process is one that doesn't have to be forced to occur after it gets started.) But returning the wheelbarrow to the top of the ramp is not spontaneous. It requires a continuous energy input. Similarly, chemical and physical changes are usually accompanied by energy changes. Some chemical reactions are spontaneous. They happen on their own. Others need a continuous energy input from an external source. Consider the reaction (Figure 1.28) of hydrogen and oxygen gases to form water vapor:



FIGURE 1.28 In the reaction shown here, a balloon was filled with appropriate amounts of hydrogen and oxygen gas. When a lit candle touched the balloon, the hydrogen and oxygen reacted explosively to form water vapor.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

This reaction is spontaneous and explosive. It releases a tremendous amount of energy. But the opposite reaction—the breakdown of water into hydrogen and oxygen gases—is not spontaneous. It occurs only if sufficient energy is continuously added, such as by passing electricity through liquid water. This process, called electrolysis, was shown in Figure 1.2, and can be described symbolically as follows:

$$2H_2O(l) \xrightarrow{\text{electrolysis}} 2H_2(g) + O_2(g)$$

But what is energy? **Energy** is the capacity to do work or to transfer heat. **Work**, usually taken to mean mechanical work, occurs when a force acts over a distance. For example, work is done when the construction worker pushes the wheelbarrow up a ramp. Work is done when compressed gases, resulting from the combustion of a fuel, push the piston in the cylinder of an automobile engine. Not all reactions can be made to do work directly, but heat energy can be harnessed to do work. For example, boiling water produces steam, which turns the turbines in power plants. The turbines spin copper coils inside a magnetic field in a generator to produce an electric current.

Energy takes many different forms, and it can be converted from one form to another. Scientists describe two types of energy: kinetic energy and potential energy. **Kinetic energy** is the energy of motion. The wheelbarrow rolling down a

ramp possesses kinetic energy. **Potential energy** is energy possessed by an object because of its position. The wheelbarrow resting at the top of the ramp has potential energy. If it did not contain this stored energy, it could not release energy when it rolled down the ramp. Any object in a position to be rolled, dropped, or otherwise allowed to move spontaneously has potential energy that will be converted to kinetic energy once the motion starts.

To distinguish kinetic from potential energy, consider the volleyball game Anna and Bill watched while they ate lunch. After they ate, Anna and Bill joined the game. When Anna served the ball, kinetic energy was transferred from her hand to the volleyball (Figure 1.29). As it ascended, the ball transferred some of its kinetic energy to the surrounding air molecules, but most of its kinetic energy was converted to potential energy. Nearly all of the kinetic energy was converted to potential energy when the ball reached the top of its ascent. As the volleyball descended, its potential energy was converted to kinetic energy. Most of its kinetic energy was transferred to the ground when Bill missed the ball. The ball's remaining kinetic energy allowed it to bounce back up, where again its kinetic energy was converted to potential energy during its ascent.

Other forms of energy—chemical, mechanical, electric, and heat energy, for example—are really just forms of kinetic or potential energy. For example, chemical compounds can release chemical energy, the energy associated with a chemical reaction. Chemical energy is potential energy arising from the positions of the atoms and molecules in the compounds. A compound releases its potential energy when it undergoes a spontaneous chemical reaction that forms substances with less potential energy. For example, the explosive material TNT (trinitrotoluene) contains considerable potential energy that is released as kinetic energy when it decomposes. Chemical compounds can also have kinetic energy. Molecules move faster as the temperature rises. The motion of molecules or atoms is associated with heat energy, or the kinetic energy that increases with increasing temperature. The fast-moving gases produced by the explosion of TNT have high kinetic energy.

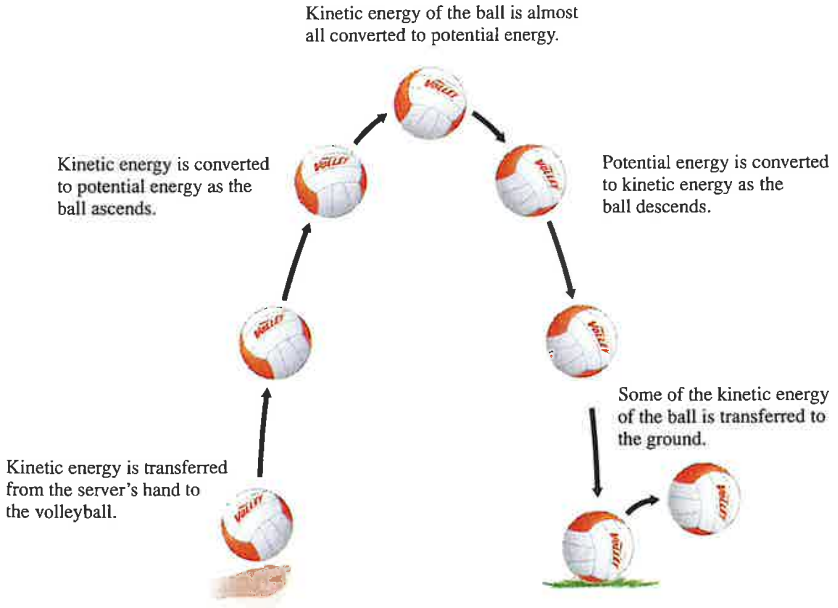
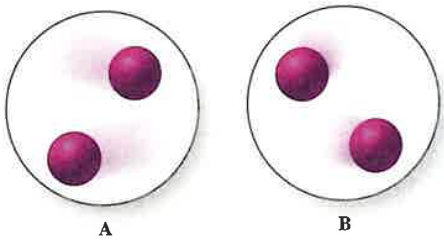


FIGURE 1.29 Kinetic energy is transferred from the server to the volleyball when it is served. As the volleyball rises in the air, its kinetic energy is converted to potential energy. When it descends, potential energy is converted to kinetic energy.

We will use the lengths of trails behind atoms or molecules to depict their relative speeds.

EXAMPLE 1.12 | Molecular Motion and Kinetic Energy

Which of these two samples of argon gas has more kinetic energy?



Solution:

The atoms that are moving faster have the greater kinetic energy. Thus, the atoms in A have the greater kinetic energy.

Practice Problem 1.12

Which of the two samples of argon gas is at a lower temperature?

Further Practice: Questions 1.101 and 1.102 at the end of the chapter

Electric energy is associated with the passage of electricity, generally through metals. Electric current passed through the filament of a lightbulb causes the metal to glow red and increases the motion of the atoms. This is a conversion of electric energy to kinetic energy. A lightbulb also gives off light energy. Nuclear energy involves both light and heat. Energy is released when one element is converted to another, as in a nuclear reactor or in the Sun, where hydrogen atoms fuse to form helium.

EXAMPLE 1.13 | Forms of Energy

Identify examples of potential and kinetic energy in this picture.



Solution:

Anything that might move in the picture has potential energy. Kinetic energy is evident in the moving people and vehicles.

Practice Problem 1.13

Identify three additional forms of energy in the photograph.

Further Practice: Questions 1.105 and 1.106 at the end of the chapter

All these forms of energy can be converted into one another. For example, the welder at the construction site starts a gasoline engine that runs a generator that makes the electricity the welder uses to join two pieces of metal together. The chemical energy in the gasoline is converted to mechanical energy that turns the generator. The mechanical energy is converted to electric energy by the generator. The electric energy is converted to heat in an arc that is formed between the welding rod and the metal to be welded. This heat melts the metal and creates the weld. Some of the electric energy is also converted to light. What energy conversions can you observe going on around you right now?

Common units of energy are calories, Calories (with a capital C), and joules. The energy input you need per day is about 2000 Calories, 2 million calories, or 8 million joules. We will study ways of measuring energy changes in Chapter 6.

1.4 | SCIENTIFIC INQUIRY

We began this chapter by describing some of the things that Anna and Bill saw around their campus. To classify the items, they observed similarities and differences in properties. They were making observations to help them understand nature.

Observation is one of the tools of scientific inquiry, but it is not the only one. The **scientific method** is an approach to asking questions and seeking answers that employs a variety of tools, techniques, and strategies. Although the scientific method is often explained as a series of steps and procedures, it is more accurately described as a way of looking at the world that differs from nonscience forms of inquiry. Scientists, like all humans, use intuition. They generalize about the world, sometimes with insufficient data. Chemists, especially, make inferences about atoms and molecules from data obtained from instruments that aren't quite capable of showing these tiny particles.

Scientists differ from professionals in nonscience disciplines in at least three important ways: (1) They test ideas by experimentation, (2) they organize their findings in particular (often mathematical) ways, (3) and they try to explain *why* things happen. Scientists use what is already known or believed about particular phenomena to gain insight into new observations from their experiments. Careful reasoning and insightful analogy are often employed, but sometimes intuition and luck play a part. Good scientists have an ability to couple objective scientific thinking with creative problem solving. In addition, a scientist must be curious enough to pursue the study of a seemingly trivial observation that can sometimes—albeit rarely—lead to a major advance in understanding.

Practicing scientists employ a variety of approaches to generate new knowledge or solve problems. Scientific inquiry generally includes observations, hypotheses, laws, and theories.

Observations

Scientific inquiry begins with ideas, knowledge, and curiosity in the minds of scientists. To look for answers to their questions, scientists collect data. Data may derive from the observation of a naturally occurring event or from deliberate experimentation. When experimenting, scientists set conditions, allow events to occur, and observe the result. This procedure permits scientists to examine events under controlled conditions not found in nature. In addition, scientists repeat one another's experiments and compare observations, thus checking the accuracy of their findings. A common experimental design involves the isolation of one factor at a time to determine which of many variables influences the outcome. The results of experiments may be qualitative (descriptive) or quantitative (numeric).

Consider this example. Suppose Anna and Bill want to experiment to find a way to clean the dull, brown coating from the pennies they saw in the fountain. After reading the results of previous studies, they think an acid might be a good agent to use. They try several, always being sure to work under the same conditions of temperature, light, ventilation, degree of tarnish, and so on. The pennies look a little



FIGURE 1.30 Copper reacts vigorously with nitric acid to form copper nitrate in solution and gaseous nitrogen dioxide.



FIGURE 1.31 When some pennies were cleaned with hydrochloric acid, one began to bubble and eventually floated.

cleaner after being placed in acetic acid, but the dull coating remains unaffected. With nitric acid, a reaction occurs. A blue-green liquid and a red gas form. The tarnish disappears, but so do the pennies (Figure 1.30). Hydrochloric acid appears to work best because the surface tarnish vanishes quickly, leaving the penny intact. However, Anna and Bill notice that one of the pennies appears to react with the hydrochloric acid to form bubbles. After a time, this penny begins to float (Figure 1.31). Upon closer examination of this penny, they notice that bubbles are coming from scratches on its surface. Their descriptions of the behavior of pennies in acids are observations made from deliberate and controlled experimentation.

Hypotheses

To organize and correlate multiple observations and sets of collected data, scientists propose hypotheses. A **hypothesis** is a tentative explanation for the properties or behavior of matter that accounts for a set of observations and can be tested. Because hypotheses are usually starting points in the explanation of natural phenomena, they normally lead to further experimentation. In practice, hypotheses are intuitive guesses that may be based on small amounts of data. Often a hypothesis is modified repeatedly in light of the results of additional experimentation. The building of scientific knowledge involves a cyclic interplay between observations and the making and testing of hypotheses (Figure 1.32).

Consider the experiments with pennies and acids. Earlier, Bill and Anna observed that one of the pennies reacted with hydrochloric acid to form a colorless gas. That gas formed along scratches on the penny's surface. Their observation might lead them to form a hypothesis: Pennies react with hydrochloric acid to form bubbles of a gas because they have scratches that make the reaction faster. How might they test this hypothesis? They could scratch several pennies, dropping them into hydrochloric acid and checking for gas bubble formation. Suppose they performed this experiment with 10 pennies (Figure 1.33) and found that 7 of them formed bubbles

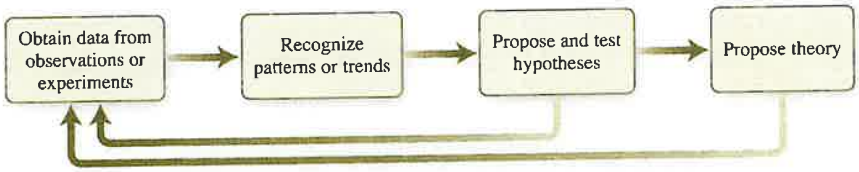


FIGURE 1.32 Observations and hypotheses are linked in a cyclic fashion during the development of scientific knowledge.



FIGURE 1.33 Some of the pennies react with hydrochloric acid at the location of the scratches, while others do not.

at the scratches, but 3 did not. They would have to conclude that their hypothesis was only partially correct. Why? Because some of the pennies did not form bubbles as predicted by their hypothesis. Therefore, they might think further, asking themselves what other characteristics of pennies could influence the outcome.

While all pennies appear the same, close observation reveals detectable differences. For example, mints located in different cities produce pennies. The first letter of the city name appears just below the year in which the penny was minted (Figure 1.34). A *D* is stamped on pennies minted in Denver; *S* is used for San Francisco; and pennies minted in Philadelphia have no identifying mark. Given this new information, Anna and Bill might state a new hypothesis: Pennies react with hydrochloric acid to form bubbles if they are scratched *and* if they come from a specific mint—perhaps because that mint formulates its pennies differently. However, examination of the pennies from the previous experiment indicates that the mint location doesn't matter. Some scratched pennies from all the mints react and some don't. The amended hypothesis is incorrect.



FIGURE 1.34 Pennies are marked with the year and city where they were minted.

As Bill and Anna set out to amend their hypothesis again, they realize they must isolate and assess one variable at a time. If they try to test too many ideas at once, they'd end up in a muddle of effects with several possible causes. So, while it's clear that scratches are necessary for bubbling to occur, some other variable or variables must be involved. As a next step, Anna and Bill reexamine the pennies from their previous experiment. They find that all the reactive pennies were minted in 1984 or more recently, while all the nonreactive pennies were minted in 1982 or before. This observation leads them to another amended hypothesis: Scratched pennies minted in 1982 or before do not react with hydrochloric acid to form bubbles, while scratched pennies minted since 1984 do—perhaps because the metal used to make the pennies was changed between 1982 and 1984.

How could they test their modified hypothesis? First, they must make sure that all old and new pennies respond in the same way as their original 10. They must then rule out any effects of chance by testing dozens, maybe even hundreds, of pennies. Assuming they get the results their hypothesis predicted, they may go on to cut some pennies open, perhaps verifying with direct observation that the insides of new and old pennies contain different materials.

Such findings would allow them to predict the behavior of pennies in hydrochloric acid, but they would not have a complete explanation for the chemical behavior. They would need to carry the process of scientific inquiry even further.

Laws

When the behavior of matter is so consistent that it appears to have universal validity, we call this behavior a *law*. A scientific law describes the way nature operates under a specified set of conditions. For example, in the late eighteenth century, observations of the amounts of materials consumed and produced in chemical reactions led to the formulation of the *law of conservation of mass*. This law states that the mass of products obtained from a chemical reaction equals the mass of the substances that react. Every known chemical reaction that has been studied follows this law. For example, we could measure the mass of a penny and the mass of the nitric acid before allowing the reaction in Figure 1.30 to proceed. Upon completion of the reaction, we would find that the mass of the blue-green liquid plus the mass of the red gas is equal to the masses of our starting materials.

When nuclear reactions were discovered—such as the fusion reactions that power our Sun—scientists realized that processes that release very large amounts of energy do not conserve mass. To accommodate this new information, the *law of conservation of mass* had to be modified to become the *law of conservation of mass and energy.*

Theories

Earlier we developed a hypothesis about the behavior of pennies in hydrochloric acid. In the Laws section we described an example of a scientific law. Hypotheses and laws only describe *how* nature works, not *why*. *Theories* explain why observations, hypotheses, or laws apply under many different circumstances. For example,



FIGURE 1.35 Zinc, but not copper, reacts with hydrochloric acid, producing hydrogen gas.

the *atomic theory*, which we will discuss in Chapter 2, explains many aspects of the behavior of matter, including the law of conservation of mass. Theories often employ mathematical or physical models that, if correct, explain the behavior of matter. Like hypotheses, a theory fits known observations. If new facts become known, theories may have to be modified or amended.

Let's return to the behavior of pennies in hydrochloric acid. Anna and Bill need to propose a theory that explains the behavior of all the pennies. Their Internet and library research reveals that before 1982 all pennies were minted from a copper alloy containing 5% zinc. After 1984 all pennies were copper-coated disks of zinc. The newer pennies contain approximately 97.5% zinc. If Anna and Bill can find out how pure zinc and pure copper react with hydrochloric acid, they might develop an explanation for the chemical behavior of pennies in the presence of hydrochloric acid. A test with pieces of copper and zinc (Figure 1.35) indicates that zinc does indeed react with hydrochloric acid to release bubbles, while copper does not. They now have an explanation for the behavior of pennies in hydrochloric acid that is consistent with all the relevant observations. They can even explain why some of the pennies float. If the zinc is completely removed by reaction with hydrochloric acid, the copper shell that remains may fill with gas and rise to the surface of the liquid.

Scientific Inquiry in Practice

There is a perception that scientists have wild hair, dress funny, wear pocket protectors, lack social graces, and work in isolation in a laboratory. This stereotype is often perpetuated in movies, but is it real? Do you know any scientists who fit this description? Perhaps you do, but the stereotype is best left to the movies. Scientists are people, and they vary as much in their appearance, personalities, and preferences as people in any other vocation (Figure 1.36). Although many scientists do indeed work independently in the lab, research groups who share common interests do most scientific work collaboratively. Sometimes they make discoveries that follow long periods of painstaking work. At other times, new insights arise rapidly through serendipity. These are accidental, fortunate breakthroughs, such as Alexander Fleming's discovery of penicillin or Henri Becquerel's discovery of radioactivity. Such serendipitous events happen only occasionally, and they happen to individuals who can recognize their importance. Both Fleming and Becquerel noticed anomalies that less skilled observers might have overlooked. As Louis Pasteur stated, "In the fields of observation, chance favors only the prepared mind."

While most of scientific inquiry proceeds through hypothesis testing, this is not the only form the scientific method takes. A new approach, especially useful in the search for new drugs, is *combinatorial chemistry*. A series of related chemical compounds is systematically prepared and tested for effectiveness in disease treatment.





FIGURE 1.36 Chemists are a diverse group of people who work in a variety of environments.

Many different combinations are tried, using techniques involving miniaturization, robotics, and computer control. The compounds are screened as possible candidates for drug action. For example, if a drug is needed that binds to a particular enzyme to produce a biological effect, then the various compounds would be added to that enzyme to see if they do bind. If so, additional testing would be carried out.

An alternate approach involves the mixing of many chemicals, producing many different products in one container. These are then tested, either in the mixture or after separation. With current techniques, it is possible for one laboratory to synthesize and test as many as 100,000 new compounds in a month.

To avoid generating a lot of chemical waste from the synthesis of compounds tested by combinatorial techniques or waste generated at the end of testing, scientists run these experiments on a microlevel scale. The machine like the one shown in Figure 1.37 can dispense as little as 10 to 50 μL in each sample well. An important initiative within the chemistry community is to develop chemical processes that prevent pollution and reduce the amount of natural resources used to manufacture various consumer, research, and industrial products. *Green chemistry* and *sustainability* are terms often given to describe these efforts.

One of the primary principles of green chemistry is to prevent waste. Industrial and research facilities are working diligently to develop processes that do not generate waste that must be treated or disposed of. One way of eliminating waste is to design products that break down to harmless materials after they are used. Another principle of green chemistry is to design processes that do not deplete natural resources. Instead, green-chemistry principles recommend using raw materials that are renewable—that is, raw materials that come from agricultural sources or from the waste products of other processes. Fossil fuels are the primary source of depleting raw materials.



FIGURE 1.37 Many chemicals can be tested at once using equipment such as that shown here. Each tip dispenses a very small quantity of chemical into a corresponding well where the chemical's action as a potential pharmaceutical agent can be tested.

SUMMARY

In this chapter we showed how things can be classified as matter or energy. We also demonstrated that matter can be classified in different ways based on its characteristics. Matter can be classified as an element, a compound, or a mixture of one or both. Elements can be further classified as metals or nonmetals, and mixtures can be further classified as homogeneous mixtures (solutions) or heterogeneous mixtures.

Elements and compounds are commonly represented using element symbols and chemical formulas. Molecules can be either elements or compounds. Models can be constructed using spheres for atoms and combinations of spheres for molecules.

Matter can also be classified by its physical state, whether solid, liquid, or gas. Each substance has a set of unique physical and chemical properties. Some physical properties, such as mass, density, or boiling point, derive from measurements using appropriate units.

A change in a substance can be classified as either physical or chemical. In a physical change, the identity of the substance(s) remains unchanged. When a new substance is formed, the change is chemical, and the process of change is often called a chemical reaction.

The energy of an object is a combination of its kinetic and potential energy. Energy can change from one form to another, and it can be transferred as heat. Energy changes, often in the form of heat, accompany chemical reactions.

The scientific method includes a variety of methods of inquiry employed by scientists. Observation, or the collection of data, is a necessary part of the scientific method. It frequently involves the design of carefully controlled experiments. Hypotheses are tentative explanations for the results of experiments. Laws result when observations appear to have universal validity, and theories are explanations for laws.

Math Toolbox 1.1 | Scientific Notation

Many numbers used in chemistry are either very large, such as 602,200,000,000,000,000,000,000 atoms in 12 g of carbon, or very small, such as 0.0000000001 centimeter per picometer (cm/pm). It is very easy to make mistakes with such numbers, so we express them in a shorthand notation called exponential or scientific notation. The numbers used in the examples are written as $(6.022 \times 10^{23}$ atoms)/(12 g C) and 1×10^{-10} cm/pm in scientific notation. In this form, the numbers are easier to write, and it is easier to keep track of the position of the decimal point when carrying out calculations.

A number in scientific notation is expressed as $C \times 10^n$, where C is the *coefficient* and n is the *exponent*. The coefficient C is a number equal to or greater than 1 and less than 10 that is obtained by moving the decimal point the appropriate number of places. The exponent n is a positive or negative integer (whole number) equal to the number of places the decimal point must be moved to give C. For numbers greater than 1, the decimal point is moved to the left and the exponent is positive:

$$523 = 5.23 \times 10^2$$

For numbers smaller than 1, the exponent is negative, since the decimal point must be moved to the right:

$$0.000523 = 5.23 \times 10^{-4}$$

EXAMPLE 1.14 Converting Between Decimal Form and Scientific Notation

(a) Convert 453,000 to scientific notation.
(b) Convert 0.00052 to scientific notation.
(c) Write the value 4.0×10^{-3} in decimal form.
(d) Write the value 8.657×10^4 in decimal form.
(e) Place the following numbers in order from smallest to largest: 10^3, 1, 0, 10^{-3}, 10^{-9}.

Solution:

(a) The number 453,000 is greater than 1, so our exponent n in 10^n will be a positive value. To convert to scientific notation, we move the decimal point between the 4 and 5. The decimal point is moved five decimal places to the left, so the exponent is 5. The value 453,000 in scientific notation is therefore 4.53×10^5.

(b) The number 0.00052 is less than 1, so our exponent n in 10^n will be a negative value. To convert to scientific notation, we move the decimal point between the 5 and 2. The decimal point is moved four decimal places to the right, so the exponent is -4. The value 0.00052 in scientific notation is therefore 5.2×10^{-4}.

(c) The number 4.0×10^{-3} has a negative exponent, so it is less than 1. To convert to decimal form, we move the decimal point to the left three places: 0.0040.

(d) The number 8.657×10^4 has a positive exponent, so it is greater than 1. To convert to decimal form, we move the decimal point to the right four places: 86,570.

(e) Zero is the lowest value since none of the numbers are negative in value. The values with negative exponents are between 0 and 1. The value with a positive exponent is the only value greater than 1. The order is $0 < 10^{-9} < 10^{-3} < 1 < 10^3$.

Practice Problem 1.14

(a) Convert 0.0123 to scientific notation.
(b) Convert 123 to scientific notation.
(c) Write the value 2.55×10^4 in decimal form.
(d) Write the value 1×10^{-3} in decimal form.
(e) Place the following numbers in order from smallest to largest: 10^{-6}, 0, 1, 10^4, -2.

Further Practice: Questions 1.3 and 1.5 at the end of the chapter

Using Your Calculator

Such operations can be carried out more easily with a scientific calculator. Methods of use vary with different brands and models, so users should always consult their manual for specific instructions. However, the general approach is as follows. First enter the coefficient, including the decimal point. Then press the appropriate button for entering the exponent followed by the value of the exponent. The button is generally labeled EE, EXP, or 10^x. If the

Math Toolbox 1.1 (*continued*)

exponent has a negative value, press the change-sign button before entering the value of the exponent. Don't confuse this button with the subtraction button. The change-sign button is usually labeled [+/−] or [(−)].

Generally, it takes fewer key strokes to input a number in exponential form using the [EE] or [EXP] buttons. For a number in the format $C \times 10^n$, the sequence of calculator buttons is C [EE] n or C [EXP] n. For example, the number 3.4×10^{-5} is input as 3.4 [EE] [+/−] 5. To test your calculator, try the mathematical operation $(8 \times 10^6) \div 2$. We input this as 8 [EE] 6 ÷ 2. The answer is 4×10^6, so 4 [EE] 6 should appear on your calculator.

Rules of Exponents

When you carry out mathematical operations with numbers written in scientific notation, you need to know how to work on the exponential part of the number. You can usually do this with your calculator, but you should always check whether an answer from a calculator makes sense. To do this, verify the calculation by hand or make an estimate in your mind. The following instructions show how to work without a calculator. You should also perform these manipulations with your calculator to assure yourself that you can do them all.

In several of the examples, the coefficient is understood to be 1. That is, 1.0×10^4 is the same as 10^4. Since 1 multiplied or divided by 1 still equals 1, the coefficient does not need to be written.

To multiply exponential notation numbers, multiply the coefficients and add the exponents:

$$10^4 \times 10^7 = 10^{(4+7)} = 10^{11}$$
$$10^3 \times 10^{-5} = 10^{(3+-5)} = 10^{-2}$$
$$10^{-4} \times 10^4 = 10^{(-4+4)} = 10^0 = 1$$
$$(2 \times 10^{-4}) \times (4 \times 10^4) = (2 \times 4) \times 10^{(-4+4)} = 8 \times 10^0 = 8$$

To divide numbers written in exponential notation, divide the coefficients and subtract the exponents:

$$\frac{10^6}{10^4} = 10^{(6-4)} = 10^2$$
$$\frac{10^{-3}}{10^{-8}} = 10^{(-3--8)} = 10^5$$
$$\frac{8 \times 10^6}{2 \times 10^4} = \frac{8}{2} \times 10^{(6-4)} = 4 \times 10^2$$

To raise an exponential number to a power, raise both the coefficient and the exponent to the power by multiplying:

$$(10^4)^3 = 10^{(4 \times 3)} = 10^{12}$$
$$(10^{-2})^{-3} = 10^{(-2 \times -3)} = 10^6$$
$$(2 \times 10^4)^3 = (2)^3 \times 10^{(4 \times 3)} = 8 \times 10^{12}$$

To extract the root of an exponential number, take the root of both the coefficient and the exponent:

$$(10^8)^{1/2} = 10^{(8/2)} = 10^4$$
$$(10^{27})^{1/9} = 10^{(27/9)} = 10^3$$
$$(4 \times 10^8)^{1/2} = (4)^{1/2} \times 10^{(8/2)} = 2 \times 10^4$$

If the exponent is not evenly divisible by the root, we rewrite the number so the exponent is evenly divisible:

$$(10^7)^{1/2} = (10 \times 10^6)^{1/2} = 10^{1/2} \times (10^6)^{1/2} = 3.16 \times 10^3$$

To add or subtract numbers in exponential notation, first express both numbers with identical exponents. To do so, shift the decimal point in one of the numbers so its exponential part is the same as that of the other number. Then simply add or subtract the coefficients and carry the exponential part through unchanged. For example, suppose we want to add the following quantities:

$$5.2 \times 10^4 + 7.0 \times 10^3$$

We first change one number so they both have the same exponents:

$$52. \times 10^3 + 7.0 \times 10^3$$

Then we can add the two coefficients:

$$(52. + 7.0) \times 10^3$$

If necessary, we shift the decimal point to return the number to proper scientific notation:

$$59. \times 10^3 = 5.9 \times 10^4$$

EXAMPLE 1.15 Mathematical Operations with Numbers Written in Scientific Notation

Perform the following operations without using your calculator and report your answers in scientific notation.

(a) $(2.0 \times 10^{-2}) \times (2.5 \times 10^4)$
(b) $(4.0 \times 10^{-2}) \div (2.0 \times 10^3)$
(c) $(3.0 \times 10^4)^2$
(d) $(4.3 \times 10^2) + (6.90 \times 10^3)$

Solution:

(a) To multiply numbers in scientific notation, multiply the coefficients and add the exponents:
$$(2.0 \times 2.5) \times (10^{-2+4}) = 5.0 \times 10^2$$

(b) To divide numbers in scientific notation, divide the coefficients and subtract the exponents:
$$(4.0 \div 2.0) \times (10^{-2-3}) = 2.0 \times 10^{-5}$$

(c) To raise a number in scientific notation to a power, raise both the coefficient and the exponent to the power:
$$(3.0 \times 10^4)^2 = (3.0)^2 \times (10^4)^2$$
When we raise an exponent to a power, we multiply the exponent by the power:
$$(3.0)^2 \times (10^4)^2 = 9.0 \times 10^8$$

(d) To add two numbers in scientific notation, we must first express both numbers with identical exponents:
$(4.3 \times 10^2) + (6.90 \times 10^3) = (0.43 \times 10^3) + (6.90 \times 10^3)$.
Then we add the coefficients:
$$(0.43 + 6.90) \times 10^3 = 7.33 \times 10^3$$

Now, confirm your answers with a calculator. If you don't get the answers shown, you're not inputting numbers in scientific notation correctly.

Practice Problem 1.15

Perform the following operations without using your calculator and report your answers in scientific notation.

(a) $(4.0 \times 10^6) \times (1.5 \times 10^{-3})$
(b) $(3.0 \times 10^8) \div (1.0 \times 10^{-7})$
(c) $(2.0 \times 10^{-2})^3$
(d) $(4.10 \times 10^{10}) - (2.0 \times 10^9)$

Further Practice: Questions 1.7 and 1.8 at the end of the chapter

Math Toolbox 1.2 | Significant Figures

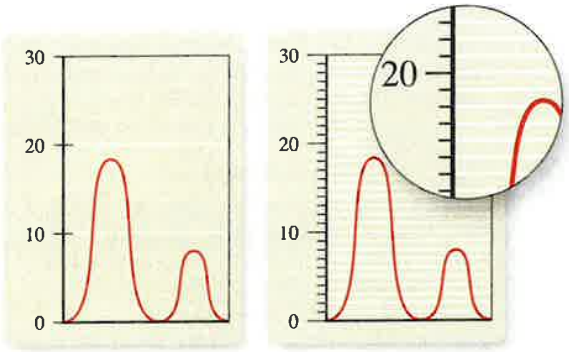
Numbers used in chemistry can be placed into two categories. Some numbers are *exact.* They are established by definition or by counting. Defined numbers are exact because they are assigned specific values: 12 in = 1 ft, 2.54 cm = 1 in, and 10 mm = 1 cm, for example. Numbers established by counting are known exactly because they can be counted with no errors. Since defined and counted numbers are known precisely, there is no uncertainty in their values.

Other numbers are not exact. They are numbers obtained by measurement or from observation. They may also include numbers resulting from a count if the number is very large. There is always some uncertainty in the value of such numbers because they depend on how closely the measuring instrument and the experimenter can measure the values.

Precision and Accuracy

Uncertainty in numbers can be described in terms of either precision or accuracy. The *precision* of a measured number is the extent of agreement between repeated measurements of its value. *Accuracy* is the difference between the value of a measured number and its expected or correct value. If repetitive measurements give values close to one another, the number is precise, whether or not it is accurate. The number is accurate only if it is close to the true value.

We usually report the precision of a number by writing an appropriate number of significant figures. *Significant figures* in a number are all the digits of which we are absolutely certain, plus one additional digit, which is somewhat uncertain. For example, if we measure the height of a line on a graph calibrated with a line every 10 cm, we might obtain a value between 10 and 20 cm, which we estimate to be 18 cm. We are certain of the 1 in 18 cm, but not of the 8. Both of these digits are considered significant; that is, they have meaning. Unless we have other information, we assume that there is an uncertainty of at least one unit in the last digit.



With a graph calibrated more finely, in centimeters, we see that the line height is indeed somewhat more than 18 cm. The line appears to reach about three-tenths of the distance between the 18 cm and 19 cm marks, so we can estimate the height at 18.3 cm. The number of significant figures is three, two of which are certain (1 and 8) and one of which is somewhat uncertain (3).

Determining the Number of Significant Figures

Every number represents a specific quantity with a particular degree of precision that depends on the manner in which it was determined. When we work with numbers, we must be able to







Math Toolbox 1.2 (*continued*)

recognize how many significant figures they contain. We do this by first remembering that nonzero digits are always significant, no matter where they occur. The only problem in counting significant figures, then, is deciding whether a zero is significant. To do so, using the following rules:

- A zero alone in front of a decimal point is not significant; it is used simply to make sure we do not overlook the decimal point (**0**.2806, **0**.002806).
- A zero to the right of the decimal point but before the first nonzero digit is simply a place marker and is not significant (0.**0**02806).
- A zero between nonzero numbers is significant (28**0**6, 0.002**0**06).
- A zero at the end of a number and to the right of the decimal point is significant (0.0028060**,** 2806.**0**).
- A zero at the end of a number and to the left of the decimal point (2806**0**) may or may not be significant. We cannot tell by looking at the number. It may be precisely known, and thus significant, or it may simply be a placeholder. If we encounter such a number, we have to make our best guess of the intended meaning.

The following table summarizes the significant figures in the numbers we just considered. The digits that are significant are highlighted.

Number	Count of Significant Figures
0.2806	4
0.002806	4
2806	4
0.0028060	5
2806.0	5
280600	4 or 5 or 6

To avoid creating an ambiguous number with zeros at its end, we write the number in scientific notation (see Math Toolbox 1.1), so that the troublesome zero occurs to the right of the decimal point. In this case, it is simple to show whether the zero is significant (2.8060×10^4) or not (2.806×10^4). The power of ten, 10^4, is not included in the count of significant figures, since it simply tells us the position of the decimal point.

EXAMPLE 1.16 Number of Significant Figures

Determine the number of significant figures in the following:

(a) 0.060520
(b) 5020.01

Solution:

(a) Zeros before the first nonzero digit are not significant. Zeros at the end of a number and to the right of the decimal point are significant. Zeros between nonzero digits are significant. Thus, all but the first two zeros are significant, giving five significant figures.
(b) Zeros between nonzero digits are significant, so all digits in this number are significant. There are six significant figures in this number.

Practice Problem 1.16

Determine the number of significant figures in the following:

(a) 14800.0
(b) 0.00076

Further Practice: Questions 1.9 and 1.10 at the end of the chapter

Significant Figures in Calculations

We must also be concerned about the proper expression of numbers calculated from measured numbers. Calculators will not do this for us, so we have to modify their output. The rules for determining the proper number of significant figures in mathematical manipulations of measured numbers are simple. To determine the proper number of significant figures in the answer to a calculation, we consider only measured numbers. We need not consider numbers that are known exactly, such as those in conversions such as 1 foot (ft) = 12 inches (in). In the following examples, we will not use units, so that we can focus on the manipulation of the numbers. Remember, however, that a unit must accompany every measured number.

Multiplication and Division

In a multiplication or division problem, the product or quotient must have the same number of significant figures as the least precise number in the problem. Consider the following example:

$$2.4 \times 1.12 = ?$$

The first number has two significant figures, and the second has three. Since the answer must match the number with the fewest significant figures, it should have two:

$$2.4 \times 1.12 = 2.7 \text{ (not } 2.69)$$

EXAMPLE 1.17 Significant Figures in Multiplication and Division

Express the answers to the following operations with the proper number of significant figures:

(a) 5.27×3.20
(b) $(1.5 \times 10^6) \times 317.832$
(c) $6.0/2.9783$
(d) $(2.01)^3$

Solution:

(a) Each number has three significant figures, so the answer should have three:

$$5.27 \times 3.20 = 16.9$$

(b) The first number has two significant figures, while the second has six. The answer should have two significant figures, the same as the number with fewer significant figures:

$$(1.5 \times 10^6) \times 317.832 = 4.8 \times 10^8$$

(c) The numerator has two significant figures; the denominator has five. The answer should have two significant figures:

$$\frac{6.0}{2.9783} = 2.0$$

Math Toolbox 1.2 (*continued*)

(d) A number raised to a power is the same as multiplying that number by itself as many times as indicated by the power, so the answer should have the same number of significant figures as the number:

$$(2.01)^3 = 2.01 \times 2.01 \times 2.01 = 8.12$$

Practice Problem 1.17

Express the answers to the following operations with the proper number of significant figures:

(a) 5.01×1.5
(b) $0.009/3.00$
(c) $3.60 \times 4.2 \times 10^{10}$
(d) $(1.85)^2$

Further Practice: Questions 1.11 and 1.12 at the end of the chapter

Addition and Subtraction

A sum or difference can only be as precise as the least precise number used in the calculation. Thus, we round off the sum or difference to the first uncertain digit. If we add 10.1 to 1.91314, for instance, we get 12.0, not 12.01314, because there is uncertainty in the tenths position of 10.1, and this uncertainty carries over to the answer. In this case, it is not the number of significant figures that is important but rather the place of the last significant digit. Consider another example:

$$0.0005032 + 1.0102 = ?$$

The first number has four significant figures and the second number has five. However, this information does not determine the number of significant figures in the answer. If we line these numbers up, we can see which digits can be added to give an answer that has significance:

$$\begin{array}{r} 0.0005032 \\ +\ 1.0102 \\ \hline 1.0107\cancel{032} \end{array}$$

The digits after the 5 in the first number have no corresponding digits in the second number, so it is not possible to add them and obtain digits that are significant in the sum. The answer is 1.0107, which has five significant figures, dictated by the position of the last significant digit in the number with the greatest uncertainty (1.0102).

EXAMPLE 1.18 Significant Figures in Addition and Subtraction

Express the answers to the following operations with the proper number of significant figures:

(a) $24 + 1.001$
(b) $24 + 1.001 + 0.0003$
(c) $428 - 0.01$
(d) $14.03 - 13.312$

Solution:

(a) The first number has no digits after the decimal point, while the second has three. The digits after the decimal point will have no significance:

$$24 + 1.001 = 25$$

(b) Only the numbers to the left of the decimal point will have significance in the answer:

$$24 + 1.001 + 0.0003 = 25$$

(c) The second number has no digits to the left of the decimal point, while the first number has all its significant digits to the left of the decimal point, so the subtracted number does not affect the answer:

$$428 - 0.01 = 428$$

(d) The first number has two digits to the right of the decimal point; the second has three. The answer should have only two digits past the decimal:

$$14.03 - 13.312 = 0.72$$

Practice Problem 1.18

Express the answers to the following operations with the proper number of significant figures:

(a) $1.003 + 1.02 + 1.0$
(b) $1.0003 - 0.10$
(c) $2.5020 - 2.50$
(d) $7.10 + 13.473$

Further Practice: Questions 1.13 and 1.14 at the end of the chapter

Rounding Off Numbers

When a number contains more digits than are allowed by the rules of significant figures, we drop the digits after the last significant figure, using the following procedure:

- If the first digit being dropped is less than 5, leave the last significant figure unchanged. (For example, rounding to three significant figures, 6.073 becomes 6.07, and 6073 becomes 6.07×10^3.)
- If the first digit being dropped is greater than 5 or is 5 followed by digits other than zero, increase the last significant figure by one unit (rounding to three significant figures, 6.077 becomes 6.08, and 60,751 becomes 6.08×10^4).
- If the first digit being dropped is 5 followed only by zeros or by no other digits, then increase the last significant figure by one unit if it is odd (6.075 becomes 6.08) but leave it unchanged if it is even (6.085 becomes 6.08).

EXAMPLE 1.19 Rounding Numbers

Round each of the following numbers to three significant figures:

(a) 3245
(b) 12.263
(c) 0.001035

Math Toolbox 1.2 (*continued*)

(d) 312,486
(e) 312,586

Solution:

(a) With four significant figures, the last digit must be dropped to convert to three significant figures. Since the second to the last digit is even, it remains unchanged: 3240 or 3.24×10^3.
(b) The last two digits must be dropped. Since the first one dropped is greater than 5, the previous digit is increased by 1: 12.3.
(c) The last digit must be dropped. Since this digit is a 5 and the previous digit is odd, the previous digit is increased by 1: 0.00104.
(d) The last three digits must be dropped. Since the first one dropped is less than 5, the previous digit is unchanged: 312,000 or 3.12×10^5.
(e) The last three digits must be dropped. Since the first one dropped is 5 followed by nonzero digits, the previous digit must be increased by 1: 313,000 or 3.13×10^5.

Practice Problem 1.19

Round each of the following numbers to two significant figures:

(a) 1.502
(b) 13.50
(c) 0.001056
(d) 1200.0
(e) 0.2035

Further Practice: Questions 1.15 and 1.16 at the end of the chapter

Math Toolbox 1.3 | Units and Conversions

All experimental sciences are based on observation. To be useful to the experimenter and to others, observations often must involve making measurements. Measurement is the determination of the size of a particular quantity—the number of nails, the mass of a brick, the length of a wall. Measurements are always defined by both a *quantity* and a *unit*, which tells what it is we are measuring.

But what system of units do we use? Most countries use the metric system, while the United States still uses primarily the English system. The English system is based on various units that are not related to one another by a consistent factor. Common units of length, for example, are inches, feet (12 in), yards (3 ft), and miles [1760 yards (yd)]. The metric system, on the other hand, uses units that are always related by a factor of 10, or by some power of 10. Units of length are centimeters (10^{-2} m), decimeters (10 cm or 10^{-1} m), meters, and kilometers (1000 m). Scientists in all countries use the metric system in their work.

The metric system is convenient to use because its units are related by powers of 10. Converting between related units is simply a matter of shifting a decimal point. The metric system adds a further convenience. It does not use arbitrary names like the English system, but rather, it defines *base units* of measure, and any multiple or fraction of these base units is defined by a special prefix.

Prefix	Factor	Symbol	Example
giga	10^9	G	1 Gg = 1,000,000,000 g
mega	10^6	M	1Mg = 1,000,000 g
kilo	10^3	k	1 kg = 1000 g
deci	10^{-1}	d	1 dg = 0.1 g
centi	10^{-2}	c	1 cg = 0.01 g
milli	10^{-3}	m	1 mg = 0.001 g
micro*	10^{-6}	μ	1 μg = 0.000001 g
nano	10^{-9}	n	1 ng = 0.000000001 g
pico	10^{-12}	p	1 pg = 0.000000000001 g

*The microgram (μg) is sometimes abbreviated mcg in the health professions.

Many science and health-related disciplines require conversions between different systems. Some of the relationships listed here may be useful in converting between English and metric units.

English Unit	Metric Unit
1 lb = 16 oz	453.6 g
1 in	2.54 cm (exactly)
1 yd	0.9144 m
1 mi	1.609 km
1 fluid oz	29.57 mL
1 qt	0.9464 L
1 gal	3.785 L
1 ft^3	28.32 L

Metric Base Units and Derived Units

In 1960 an organization of scientists met in France to determine standards for scientific measurements. This group established the SI units (from the French, *Système Internationale*) listed here.

Unit	Symbol	Quantity
meter	m	length
kilogram	kg	mass
second	s	time
ampere	A	electric current
kelvin	K	temperature
mole	mol	amount of substance
candela	cd	luminous intensity

All other units, called *derived units,* are based on these by multiplication or division and then applying the appropriate prefix or combination of units. For example, if the mass of an object is measured in grams, it could be converted to the SI unit of mass, the kilogram, by dividing the measured mass in grams by 1000 and

Math Toolbox 1.3 (*continued*)

adding the prefix *kilo* to the quantity. (We'll discuss conversions of units in the next section.) An example of a derived unit that involves a combination of units is volume. The SI derived unit for volume is cubic meters (m^3), which for a box would mean measuring length, width, and height in meters, and then multiplying them together. Another derived unit, discussed in Section 1.2, is density. Density is a ratio of mass to volume for a given substance. What is the SI derived unit for density?

Conversion of Units

Measurements must be interpreted, often by mathematical manipulation of the data. This manipulation often involves converting one set of units into another using relationships between them. Two approaches to such conversions are *ratios* and *dimensional analysis*. In both cases, an analysis of the units provides clues to the correct solution of the problem.

Consider a problem that can be solved almost automatically: "Convert 30 min into hours." You probably gave the answer without thinking: 30 min is one-half hour (0.50 h). But how did you know this answer? Let's systematize the problem-solving process. First—even if we don't always consciously think about it—we must decide what the problem is asking for. This problem requires converting a number of minutes into the corresponding number of hours. Next, we must know the number of minutes in 1 h. From this information, write a mathematical expression that shows the equivalence of two quantities having different units:

$$1\text{ h} = 60\text{ min}$$

We can use this expression in two ways to solve this problem. First, we will describe the ratio approach.

Ratio Approach

In this approach, we use a known relationship to compare with our unknown relationship. We know how many minutes make up 1 h, but we want to know how many hours there are in 30 min. From the known relationship, we develop a ratio of equivalent quantities that is equal to 1 and that has different units in the numerator and denominator. To develop a ratio to convert minutes into hours, we use the equivalent quantities and divide 1 h by 60 min.

$$\frac{1\text{ h}}{60\text{ min}}$$

The ratio for our unknown relationship should equal the known ratio as long as we use the same units, so we can set the unknown ratio equal to the known ratio:

$$\frac{x\text{ h}}{30\text{ min}} = \frac{1\text{ h}}{60\text{ min}}$$

We can cross multiply to solve for x:

$$x\text{ h} \times 60\text{ min} = 1\text{ h} \times 30\text{ min}$$

We can then solve for x by dividing both sides by 60 min:

$$\frac{x\text{ h} \times \cancel{60\text{ min}}}{\cancel{60\text{ min}}} = \frac{1\text{ h} \times 30\ \cancel{\text{min}}}{60\ \cancel{\text{min}}}$$

We then carry out the mathematical operations:

$$x\text{ h} = \frac{1\text{ h} \times 30\ \cancel{\text{min}}}{60\ \cancel{\text{min}}} = 0.5\text{ h}$$

Notice that the units of minutes cancel out, leaving the units of hours, which is the unit called for in the problem.

Dimensional Analysis

We also use the ratio of equivalent quantities in the dimensional-analysis approach. In this approach, however, we get to the form of the manipulated ratio more quickly. We multiply the known quantity by the ratio, so that units cancel and we get the unknown quantity with the desired units:

$$\text{Time in h} = 30\text{ min} \times \frac{1\text{ h}}{60\text{ min}} = \frac{30 \times 1 \times \cancel{\text{min}} \times \text{h}}{60\ \cancel{\text{min}}} = 0.5\text{ h}$$

Notice how the starting units (minutes) cancel out, leaving the desired units (hours) to accompany the calculated number. In the dimensional-analysis approach, units in the numerator and denominator of a fraction are treated exactly the same as numbers—canceled out, multiplied, divided, squared, or whatever the mathematical operations demand. We set up the conversions so that desired units are introduced and beginning units cancel out.

Now reexamine the conversion just developed. From the same equivalence expression, 1 h = 60 min, we can derive another ratio by dividing 60 min by 1 h:

$$\frac{60\text{ min}}{1\text{ h}}$$

What happens when the original quantity is multiplied by this conversion factor?

$$\text{Time in h} = 30\text{ min} \times \frac{60\text{ min}}{1\text{ h}} = \frac{30 \times 60 \times \text{min} \times \text{min}}{1\text{ h}}$$

$$= 1800\ \frac{\text{min}^2}{\text{h}}$$

This answer is *clearly wrong* because the units make no sense. There are always two possible ratios of the known equivalent quantities, but only one gives the proper units in the answer. Which ratio is appropriate to use depends on how the problem is stated. But in every case, using the proper ratio cancels out the old units and leaves intact the desired units. Using the improper ratio results in nonsensical units (like min^2/h). This is a clear warning that the ratio was upside down or otherwise incorrect.

EXAMPLE 1.20 Converting Units

Use the dimensional-analysis approach to convert 75.0 km to meters.

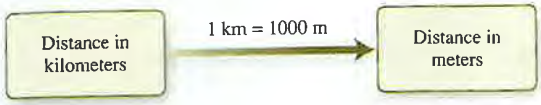
Solution:

We begin by determining if there is a relationship between kilometers and meters:

Math Toolbox 1.3 (*continued*)



From the table of metric units we see that 1 km = 1000 m, so there is a relationship between these quantities:



From the equivalence expression, 1 km = 1000 m, we can derive two ratios (conversion factors), one of which we'll use in the calculation:

$$\frac{1\text{ km}}{1000\text{ m}} \text{ or } \frac{1000\text{ m}}{1\text{ km}}$$

Only one of these expressions, when multiplied by the given quantity in kilometers, will correctly cancel the units. What are the units if we multiply 75.0 km by the first ratio?

$$75.0\text{ km} \times \frac{1\text{ km}}{1000\text{ m}} = 0.0750\,\frac{\text{km}^2}{\text{m}}$$

The units tell us that our first attempt leads to an answer that is nonsense. Applying the other conversion factor, units cancel and lead to an answer that makes sense:

$$75.0\,\cancel{\text{km}} \times \frac{1000\text{ m}}{1\,\cancel{\text{km}}} = 75{,}000\text{ m}$$

Notice that our given quantity has three significant figures. To show that we know the answer in the conversion to three significant figures, we should express it in scientific notation: 7.50×10^4 m. (See Math Toolboxes 1.1 and 1.2 for details about scientific notation and significant figures.)

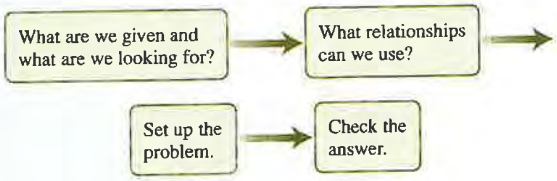
Practice Problem 1.20

Convert 0.0276 kg to grams.

Further Practice: Questions 1.17 and 1.18 at the end of the chapter

An Approach to Problem Solving

The processes just discussed can be generalized by the following diagram:



Consider the following problem: A liquid fertilizer tank has a volume of 6255 ft^3. How many gallons of liquid fertilizer can fit into this tank?

1. Decide what the problem is asking for.

First, read the problem carefully. If you are not sure what a term means, look it up. If it's necessary to use an equation to solve the problem, be sure you understand the meaning of each symbol. Look for clues in the problem itself—words or phrases such as *determine, calculate, what mass, what volume,* or *how much.* After deciding what quantity the problem is asking for, write the units in which this quantity must be stated. In the fertilizer tank example, we want to find the volume of the tank in units of gallons from units of cubic feet.

2. Decide what relationships exist between the information given in the problem and the desired quantity.

If necessary, recall or look up equivalence relationships (such as 1 h = 60 min) inside the back cover. This information may not be given in the problem itself. In the fertilizer tank example, we need to convert volume in cubic feet to volume in gallons, so we need an equivalence relationship between those two quantities. The expression is 1 ft^3 = 7.481 gal.

The additional information that describes the relationship between two quantities won't always be a direct equivalence. A series of equivalence expressions and their derived conversion factors may be needed. Sometimes a mathematical equation is required to express the equivalence between two quantities [e.g., $T_{°C} = (T_{°F} - 32)/1.8$]. At other times, some chemical principle may have to be applied.

A word of caution: Just as a problem may contain less information than you need to solve it, it may also contain more information than you need. Never assume that you must use all of the information. Examine all information critically and reject any that is not pertinent.

3. Set up the problem logically, using the relationships decided upon in step 2.

Starting with the relationships you obtained in step 2, develop the ratios needed to arrive at the final answer. Be sure to set up the ratios so that the old units cancel out and the desired units are introduced in the appropriate positions.

In the fertilizer tank example, the ratio is (7.481 gal)/(1 ft^3). The conversion is as follows:

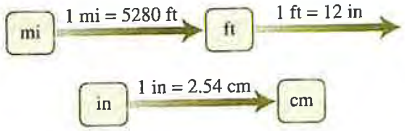
$$\text{Volume in gal} = 6255\,\cancel{\text{ft}^3} \times \frac{7.481\text{ gal}}{1\,\cancel{\text{ft}^3}} = 46{,}790\text{ gal}$$

If you need a series of ratios to find the final answer, it is often helpful to "map out" the route you will follow to get there. Suppose you need to know how many centimeters are in 1.00 mile (mi). This conversion can be summarized by a diagram:



Math Toolbox 1.3 (*continued*)

Unfortunately you can only find a table that gives the number of feet in a mile, the number of inches in a foot, and the number of centimeters in an inch. The diagram to get from miles to centimeters might look like this:



The problem setup would follow this progression:

$$\text{Distance in cm} = 1.00 \cancel{\text{mi}} \times \frac{5280 \cancel{\text{ft}}}{1 \cancel{\text{mi}}} \times \frac{12 \cancel{\text{in}}}{1 \cancel{\text{ft}}} \times \frac{2.54 \text{ cm}}{1 \cancel{\text{in}}}$$

$$= 1.61 \times 10^5 \text{ cm}$$

4. Check the answer to make sure it makes sense, both in magnitude and in units.

This step is just as important as the others. You must develop an intuitive feeling for the correct magnitude of physical quantities. Suppose, for example, you were calculating the volume of liquid antacid an ulcer patient needed to neutralize excess stomach acid, and the result of your calculation was 32 L. Now, units of liters looks reasonable, since the liter is a valid unit of volume; but it should be obvious that something is wrong with the numerical answer. Since a liter is about the same volume as a quart, this patient would need to drink 8 gal of antacid! The source of the problem is probably an arithmetic error, or possibly omission of the metric prefix *milli-* (10^{-3}) somewhere, since 32 mL would be an appropriate volume to swallow. In any event, a result such as this one should be recognized as unacceptable and the setup and calculations checked for errors.

In the fertilizer tank example, the units cancel properly, so the answer should be correct, if there are no arithmetic errors. Since there are about 7.5 gal in 1 ft^3, the answer should be greater than the volume in cubic feet by something less than a factor of 10. The answer is indeed within this expected range.

EXAMPLE 1.21 Problem Solving

If a laser beam fired from the moon takes 1.30 s to reach Earth, what is the distance in meters between the moon and Earth? Light travels in a vacuum at a speed of 3.00×10^{10} cm in 1 s or 3.00×10^{10} cm/s.

Solution:

We begin by determining if there is a relationship between seconds and meters:



There isn't, but the speed given, 3.00×10^{10} cm/s, is a relationship between time and distance in centimeters. We could use this relationship to calculate the distance in centimeters and then convert to meters.



The ratios for converting seconds to centimeters can be expressed in two ways:

$$\frac{3.00 \times 10^{10} \text{ cm}}{1 \text{ s}} \text{ or } \frac{1 \text{ s}}{3.00 \times 10^{10} \text{ cm}}$$

When multiplied by the time, 1.30 s, the first one cancels units of seconds and gives an answer in centimeters. To convert centimeters to meters, we use the equivalence expression

$$1 \text{ cm} = 10^{-2} \text{ m}$$

Rather than working with negative exponents, it is sometimes useful to work in whole numbers. Dividing both sides of the expression by the quantity of the negative exponent provides a whole number ratio that may be more convenient:

$$\frac{1 \text{ cm}}{10^{-2}} = \frac{\cancel{10^{-2}} \text{ m}}{\cancel{10^{-2}}}$$

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

Two ratios can be written to describe this relationship:

$$\frac{100 \text{ cm}}{1 \text{ m}} \text{ or } \frac{1 \text{ m}}{100 \text{ cm}}$$

When multiplied by the distance in centimeters. the second expression will correctly cancel units and give an answer in meters.

Multiplying the given distance by the two conversion factors provides a sequence of steps that allows us to cancel the time units (seconds) and distance in centimeters, giving the distance from the moon to Earth in meters:

$$\text{Distance in meters} = 1.30 \cancel{\text{s}} \times \frac{3.00 \times 10^{10} \cancel{\text{cm}}}{1 \cancel{\text{s}}} \times \frac{1 \text{ m}}{100 \cancel{\text{cm}}}$$

$$= 3.90 \times 10^8 \text{ m}$$

Does this answer make sense? Yes, because we would expect there to be a great distance between the moon and Earth. If we had inversed the conversion factors, we would have gotten 4.33×10^{-9} and units that do not make sense. The answer, 3.90×10^8 m, has three significant figures, which is the same as the number in the given quantity.

Practice Problem 1.21

A tablet of a typical pain reliever contains 200 mg of active ingredient. How many pounds of active ingredient are in a tablet? (Assume there are three significant figures in the 200-mg quantity.)

Further Practice: Questions 1.19 and 1.20 at the end of the chapter

KEY RELATIONSHIPS

Relationship	Equation
The density of an object is the ratio of its mass and volume.	Density = mass/volume
The absolute temperature in kelvins is offset from the Celsius temperature by 273.15.	$T_K = T_{°C} + 273.15$
The Fahrenheit temperature has degrees that are 1.8 times as large as Celsius degrees. The freezing point of water is set at 32°F and 0°C.	$T_{°F} = 1.8(T_{°C}) + 32$

KEY TERMS

aqueous solution (1.1)
atom (1.1)
chemical change (1.2)
chemical formula (1.1)
chemical property (1.2)
chemical reaction (1.2)
compound (1.1)
density (1.2)
element (1.1)
element symbol (1.1)
energy (1.3)
gas (1.1)
heterogeneous mixture (1.1)
homogeneous mixture (1.1)
hypothesis (1.4)
kinetic energy (1.3)
liquid (1.1)
mass (1.1)
matter (1.1)
metal (1.1)
mixture (1.1)
molecule (1.1)
nonmetal (1.1)
physical change (1.2)
physical property (1.2)
physical state (1.1)
potential energy (1.3)
pure substance (1.1)
scientific method (1.4)
solid (1.1)
solution (1.1)
temperature (1.2)
volume (1.2)
work (1.3)

QUESTIONS AND PROBLEMS

The following questions and problems, except for those in the *Additional Questions* section are paired. Questions in a pair focus on the same concept. Answers to the odd-numbered questions and problems are in Appendix D.

Matching Definitions with Key Terms

1.1 Match the key terms with the following descriptions.
(a) a measure of the quantity of matter
(b) a characteristic of a substance involving the possible transformations that the substance can undergo to produce a new substance
(c) a combination of two or more substances that can be separated by physical means
(d) a pure substance that cannot be broken down into simpler stable substances in a chemical reaction
(e) the capacity to do work or to transfer heat
(f) a characteristic of a substance that can be observed without changing its composition
(g) the physical state in which matter has no characteristic shape but takes the shape of the filled portion of its container
(h) the ratio of the mass of a substance to its volume
(i) a mixture with uniform composition
(j) the physical state of matter characterized by a fixed shape and low compressibility

1.2 Match the key terms with the following descriptions.
(a) the smallest particle of an element that retains the characteristic chemical properties of that element
(b) a change in which substances are converted into new substances that have compositions and properties different from those of the original substances
(c) anything that occupies space and is perceptible to the senses
(d) a substance composed of two or more elements combined in definite proportions
(e) a combination of atoms of one or more elements
(f) a process characterized by changes only in the physical properties of a substance, not in its composition
(g) the physical state in which matter has no fixed shape or volume but expands to fill its container completely
(h) the energy possessed by an object because of its position
(i) a tentative explanation for the properties or behavior of matter that accounts for a set of observations and can be tested
(j) the energy possessed by an object because of its motion

Math Toolbox Questions

1.3 Convert each of the following values to scientific notation.
(a) 29,500
(b) 0.000082
(c) 650,000,000
(d) 0.0100

1.4 Convert each of the following values to scientific notation.
(a) 0.00010
(b) 4500
(c) 90,100,000
(d) 0.0000079

1.5 Convert each of the following values from scientific notation to decimal form.
(a) 1.86×10^{-5} (c) 4.53×10^{5}
(b) 1×10^{7} (d) 6.1×10^{-3}

1.6 Convert each of the following values from scientific notation to decimal form.
(a) 8.2×10^{3} (c) 7×10^{-2}
(b) 2.025×10^{-6} (d) 3.0×10^{8}

1.7 For each of the following, carry out the mathematical operation and report answers in scientific notation.
(a) $(3.1 \times 10^{5}) \times (2.0 \times 10^{-2})$
(b) $(7.0 \times 10^{9}) \div (2.0 \times 10^{2})$
(c) $(2.8 \times 10^{-4}) \div (9.6 \times 10^{-2})$
(d) $(5.0 \times 10^{-4})^{2}$
(e) $(8.50 \times 10^{5}) - (3.0 \times 10^{4})$
(f) $(6.4 \times 10^{-3}) \div (4.0 \times 10^{3})$

1.8 For each of the following, carry out the mathematical operation and report answers in scientific notation.
(a) $(8.2 \times 10^{6}) \div (4.1 \times 10^{-2})$
(b) $(3.5 \times 10^{3})^{4}$
(c) $(6.6 \times 10^{-5}) \div (2.2 \times 10^{5})$
(d) $(4.23 \times 10^{2}) \times (2.0 \times 10^{-8})$
(e) $(8.44 \times 10^{5}) + (1.2 \times 10^{4})$
(f) $(4.58 \times 10^{5}) \div (3.64 \times 10^{-3})$

1.9 Determine the number of significant figures in the following.
(a) 0.0950 (d) 0.0052
(b) 760 (e) 3.46×10^{15}
(c) 1.005

1.10 Determine the number of significant figures in the following.
(a) 1.48×10^{-2} (d) 70.25
(b) 156.0 (e) 140
(c) 0.04350

1.11 Express the answers to the following operations with the proper number of significant figures.
(a) 1.2×1.216
(b) 3.000/2.0
(c) $(1.201 \times 10^{3}) \times (1.2 \times 10^{-2})$
(d) $(1.44)^{1/2}$

1.12 Express the answers to the following operations with the proper number of significant figures.
(a) $(1.600 \times 10^{-7})(2.1 \times 10^{3})$
(b) $(1.33)^{3}$
(c) 1.93×2.651
(d) 4.4/2.200

1.13 Express the answers to the following operations with the proper number of significant figures.
(a) 1.6 + 1.15 (c) 1.8 + 1.022 + 0.001
(b) 41.314 – 41.03 (d) 0.050 – 0.0012

1.14 Express the answers to the following operations with the proper number of significant figures.
(a) 87.5 + 1.3218 (c) 1.2316 – 1.01
(b) 422 – 410.2 (d) 1.5678 + 0.23 + 0.0006

1.15 Round each of the following numbers to three significant figures.
(a) 1.2126 (d) 42.186
(b) 0.2045 (e) 0.007101
(c) 1.8351

1.16 Round each of the following numbers to three significant figures.
(a) 0.020450 (d) 16.225
(b) 1.3602×10^{4} (e) 1.0001
(c) 13.475

1.17 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 36 mm to m (g) 168 lb to g
(b) 357 kg to g (h) 914 qt to L
(c) 76.50 mL to L (i) 44.5 cm to in
(d) 0.0084670 m to cm (j) 236.504 g to lb
(e) 597 nm to m (k) 2.0 L to qt
(f) 36.5 in to cm

1.18 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 75.5 km to m (g) 25.6 lb to g
(b) 25.7 g to mg (h) 4.005 qt to L
(c) 0.516 L to dL (i) 934 cm to in
(d) 5.2 cm to m (j) 155 g to lb
(e) 0.000000450 m to nm (k) 22.4 L to qt
(f) 12 in to cm

1.19 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 947 m to mi (g) 689 lb to kg
(b) 6.74 kg to lb (h) 0.5 qt to mL
(c) 250.4 mL to gal (i) 125 cm to ft
(d) 2.30 dL to mL (j) 542 mg to lb
(e) 0.000450 cm to nm (k) 25 nL to gal
(f) 37.5 in to m

1.20 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 32 cm to ft (g) 2.7 lb to kg
(b) 0.579 kg to lb (h) 8.320 qt to mL
(c) 22.70 μL to qt (i) 375 km to ft
(d) 9212 mm to km (j) 62 g to oz
(e) 465 nm to mm (k) 752 mL to gal
(f) 4 ft to cm

1.21 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 375 m/s to ft/min
(b) 24.5 cm^3 to in^3
(c) 19.3 g/mL to lb/in^3

1.22 Carry out the following conversions. Report your answers to the correct number of significant figures.
(a) 27 ft/s to cm/min
(b) 2764 ft^3 to m^3
(c) 0.927 g/mL to lb/gal

Matter and Its Classification

1.23 How would you classify the following items observed by Bill and Anna?
(a) water in a fountain containing dissolved dye
(b) a copper pipe
(c) the contents of a balloon after blowing it up by mouth
(d) a slice of pizza

1.24 How would you classify the following items observed by Bill and Anna?
(a) sand in a volleyball court
(b) a baseball bat made entirely of aluminum
(c) the contents of a balloon filled from a helium tank
(d) a glass filled with a soft drink

1.25 Which of the following are examples of matter?
(a) sunlight
(b) gasoline
(c) automobile exhaust
(d) oxygen gas
(e) iron pipe

1.26 Which of the following are not examples of matter?
(a) light from a fluorescent bulb
(b) sand
(c) wheelbarrow rolling down a ramp
(d) helium balloons
(e) heat from a welding torch

1.27 How are elements distinguished from compounds?

1.28 How are homogeneous mixtures distinguished from heterogeneous mixtures?

1.29 List characteristics of metals.

1.30 List characteristics of nonmetals.

1.31 Name the following elements.
(a) Ti (b) Ta (c) Th (d) Tc (e) Tl

1.32 Name the following elements.
(a) C (e) Cu
(b) Ca (f) Cl
(c) Cr (g) Cs
(d) Co

1.33 Name the following elements.
(a) B (b) Ba (c) Be (d) Br (e) Bi

1.34 Name the following elements.
(a) S (b) Si (c) Se (d) Sr (e) Sn

1.35 Name the following elements.
(a) N (c) Mn (e) Al
(b) Fe (d) Mg (f) Cl

1.36 Name the following elements.
(a) Be (c) Ni (e) Ti
(b) Rb (d) Sc (f) Ne

1.37 What are the symbols for the following elements?
(a) iron Fe (d) gold – Au
(b) lead Pb (e) antimony – Sb
(c) silver Ag

1.38 What are the symbols for the following elements?
(a) copper (d) sodium
(b) mercury (e) tungsten
(c) tin

1.39 A chemical novice used the symbol Ir to represent iron. Is this an acceptable symbol for the element? If not, what is the correct symbol for iron?

1.40 A chemical novice used the symbol SI to represent silicon. Is this an acceptable symbol for the element? If not, what is the correct symbol?

1.41 The symbol NO was used by a student to represent nobelium, an unstable, synthetic element. Is this an acceptable symbol for the element? If not, what is the correct symbol?

1.42 A student used the symbol CO to represent cobalt, an element found in vitamin B_{12}. Is this an acceptable symbol for the element? If not, what is the correct symbol?

1.43 Classify each of the following as a pure substance, a homogeneous mixture (solution), or a heterogeneous mixture: hamburger, salt, soft drink, and ketchup.



1.44 Classify each of the following as a pure substance, a homogeneous mixture (solution), or a heterogeneous mixture: sand, boardwalk, ocean, and roller coaster.



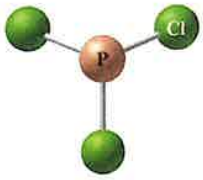
1.45 Elemental hydrogen normally exists as two hydrogen atoms bound together. Write the formula for this molecule, and draw a picture to represent how it might look on a molecular level.

1.46 Elemental chlorine normally exists as two chlorine atoms bound together. Write the formula for this molecule, and draw a picture to represent how it might look on a molecular level.

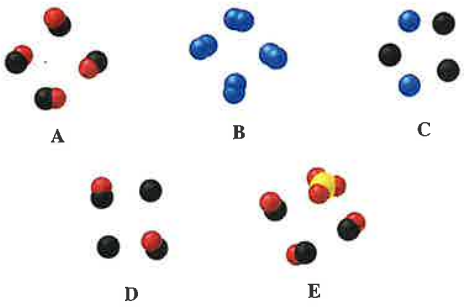
1.47 This image is a representation for a compound containing nitrogen and oxygen. Write the formula for this compound.



1.48 This image represents a compound containing phosphorus and chlorine. Write the formula for this compound.



1.49 Which of the images represents a mixture of an element and a compound?



1.50 Which of the images in Question 1.49 represents a pure substance that is a compound?

1.51 Classify each of the following as an element or a compound.

(a) O_2 (c) P_4 (e) NaCl
(b) Fe_2O_3 (d) He (f) H_2O

1.52 Classify each of the following as an element or a compound.

(a) hydrogen gas (d) nitrogen dioxide
(b) water (e) aluminum chloride
(c) salt (f) neon

1.53 Under normal conditions, mercury is a liquid. Draw molecular-level pictures of what mercury atoms might look like in the liquid and solid states.

1.54 Under normal conditions, bromine is a liquid. Draw molecular-level pictures of what bromine, Br_2, might look like in the liquid and gaseous states.

1.55 What type of matter expands to fill its container and can be compressed to a smaller volume?

1.56 What type of matter is composed of particles that do not move past one another?

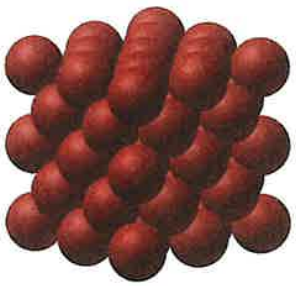
1.57 Identify the physical state of each of the following elements from their symbols.

(a) $Cl_2(g)$ (b) $Hg(l)$ (c) $C(s)$

1.58 Identify the physical state of each of the following compounds from their symbols.

(a) $NaCl(s)$ (b) $CH_3OH(l)$ (c) $CO_2(g)$

1.59 What physical state is represented in this diagram?



1.60 Draw a picture of the gaseous state of the substance shown in Question 1.59.

1.61 How might you symbolically represent a homogeneous mixture of oxygen, O_2, and water?

1.62 Why does the symbol $H_2O(aq)$ make no sense?

Physical and Chemical Changes and Properties of Matter

1.63 At the beginning of the chapter, Anna and Bill made many observations about what they saw. Some of their observations included color, texture, and shininess. Are these physical or chemical properties?

1.64 At the beginning of the chapter, you were asked to classify processes like a truck running on gas or welding pieces of metal. After reading the chapter, would you classify these processes as physical or chemical changes?

1.65 A slice of Swiss cheese contains 45 mg of sodium.

(a) What is this mass in units of grams?
(b) What is this mass in units of ounces (oz)? (16 oz = 453.6 g)
(c) What is this mass in pounds (lb)? (1 lb = 453.6 g)

1.66 A package of Swiss cheese has a mass of 0.340 kg.

(a) What is this mass in grams?
(b) What is this mass in ounces (oz)? (16 oz = 453.6 g)
(c) What is this mass in pounds (lb)? (1 lb = 453.6 g)

1.67 A grain of salt has a mass of about 1.0×10^{-4} g. What is its mass in the following units?

(a) milligrams
(b) micrograms
(c) kilograms

1.68 If a dog has a mass of 15.2 kg, what is its mass in the following units?

(a) grams
(b) milligrams
(c) micrograms

1.69 If you drank 1.2 L of a sports drink, what volume did you consume in the following units?

(a) milliliters
(b) cubic centimeters
(c) cubic meters

1.70 If the volume of helium in a balloon is 145 cm^3, what is its volume in the following units?

(a) milliliters
(b) liters
(c) cubic meters

1.71 If the length, width, and height of a box are 8.0 cm, 5.0 cm, and 4.0 cm, respectively, what is the volume of the box in units of milliliters and liters?

1.72 If a cubic box (all sides the same length) has a volume of 1.0 L, what is the length of each side of the box?

1.73 A slice of cheese has a mass of 28 g and a volume of 21 cm^3. What is the density of the cheese in units of g/cm^3 and g/mL?

1.74 Two stones resembling diamonds are suspected of being fakes. To determine if the stones might be real, the mass and volume of each were measured. Both stones have the same volume, 0.15 cm^3. However, stone A has a mass of 0.52 g and stone B has a mass of 0.42 g. If diamond has a density of 3.5 g/cm^3, could the stones be real diamonds? Explain.

1.75 If the density of a sugar solution is 1.30 g/mL, what volume of this solution has a mass of 50.0 g?

1.76 The density of a certain type of plastic is 0.75 g/cm^3. If a sheet of this plastic is 10.0 m long, 1.0 m wide, and 1 cm thick, what is its mass?

1.77 Why do liquids have greater densities than gases?

1.78 When a balloon filled with air is heated, the balloon increases in volume. Does the density of the air in the balloon increase, decrease, or remain the same?

1.79 A piece of plastic sinks in oil but floats in water. Place these three substances in order from lowest density to greatest density.

1.80 What special molecular-level feature of ice explains why ice floats in water?



Ice



Liquid water

1.81 Acetone, a component of some types of fingernail polish, has a boiling point of 56°C. What is its boiling point in units of kelvin?

1.82 The boiling point of liquid nitrogen is 77 K. What is its boiling point in units of degrees Celsius?

1.83 What is the difference in temperature between the boiling point of water and the freezing point of water in each of the following temperature scales?
(a) Celsius scale
(b) kelvin scale
(c) Fahrenheit scale

1.84 If the temperature of a cup of coffee decreases from 60.0°C to 25.0°C, what is the decrease in temperature in units of degrees Celsius and kelvin?

1.85 Does the boiling point of a substance depend on how much of this substance you have?

1.86 Does the melting point of a substance depend on how much of this substance you have?

1.87 Identify each of the following as a physical property or a chemical property.
(a) mass
(b) density
(c) flammability
(d) resistance to corrosion
(e) melting point
(f) reactivity with water

1.88 Identify each of the following as a physical property or a chemical property.
(a) boiling point
(b) reactivity with oxygen
(c) resistance to forming compounds with other elements
(d) volume

1.89 Identify each of the following as a physical change or a chemical change.
(a) boiling acetone
(b) dissolving oxygen gas in water
(c) combining hydrogen and oxygen gas to make water
(d) burning gasoline
(e) screening rocks from sand
(f) the conversion of ozone to oxygen,
$2O_3(g) \longrightarrow 3O_2(g)$

1.90 Identify each of the following as a physical change or a chemical change.
(a) condensation of ethanol
(b) combining zinc and oxygen to make the compound zinc oxide
(c) dissolving sugar in water
(d) burning a piece of paper
(e) combining sodium metal with water, producing sodium hydroxide and hydrogen gas
(f) filtering algae from water

1.91 Write a symbolic representation and a molecular-level representation for the change that occurs during the condensation of chlorine, Cl_2.

1.92 Write a symbolic representation and a molecular-level representation for the process of freezing oxygen, O_2.

1.93 Do the changes shown in this diagram represent a physical or chemical change?



Before



After

1.94 Do the changes shown in this diagram represent a physical or chemical change?



Before

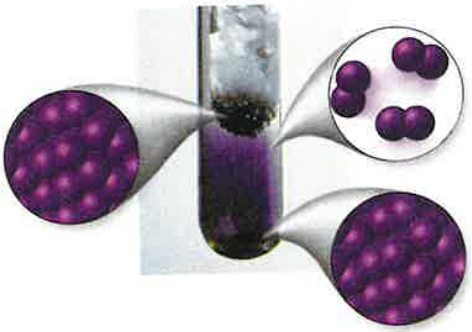


After

1.95 Draw a picture that shows CH_4 (shown in the "After" image in Question 1.93) condensing from a gas to a liquid. Does this picture represent a physical or a chemical change?

1.96 Draw a picture that shows water boiling. Does this picture represent a physical or chemical change?

1.97 The image shows what happens when iodine, I_2, is placed in the bottom of a test tube and heated. Is this a physical or chemical change?



1.98 The picture shows natural gas (CH_4) burning. In the process, carbon dioxide and water form. Is this a physical change or a chemical change?



Energy and Energy Changes

1.99 Anna and Bill saw a construction worker welding pipe. Classify the forms of energy they were observing.

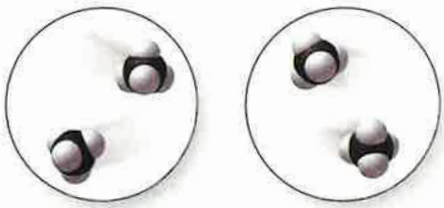
1.100 Bill and Anna watched students playing volleyball in the sunshine. Classify the forms of energy they were observing.

1.101 Which of these two samples of carbon dioxide gas has more kinetic energy? Explain your answer.



A B

1.102 Which of these two samples of methane gas is at a higher temperature? Explain your answer.



A B

1.103 Give examples of potential energy that you can find in your room.

1.104 Give examples of kinetic energy that you can find in your room.

1.105 Distinguish between the different types of energy shown in the following picture.



1.106 Distinguish between the different types of energy shown in the following picture.



1.107 Describe how some common types of energy can change from one type to another.

1.108 If energy cannot be created, what is the source of the energy that is released when gasoline is burned?

1.109 In terms of kinetic and potential energy, describe what happens when a home run is hit on a softball field.

1.110 In terms of kinetic and potential energy, describe what happens when a basketball is dribbled down a basketball court and shot to score three points.

1.111 Body mass index (BMI) is a number calculated from a person's weight and height. BMI provides a reliable indicator of body fatness for most people and is used to screen for weight categories that may lead to health problems. The formula for a person's BMI is

$$\text{BMI} = \frac{\text{weight (kg)}}{[\text{height (m)}]^2}$$

For an adult, a BM value between 18.5 and 24.9 is healthy and considered normal.

(a) Calculate the BMI for a 6-ft, 2-in man who weighs 169 lb.

(b) Is this person underweight, at a healthy weight, or overweight?

1.112 Consider propane fuel being burned to power a vehicle. How would you classify the transformation of energy in this process in terms of kinetic and potential energy?

1.113 Consider water falling into a fountain. How would you classify the transformation of energy in this process in terms of kinetic and potential energy?

Scientific Inquiry

1.114 Explain the difference between a hypothesis and a theory.

1.115 Explain how a hypothesis is used in scientific research.

1.116 Classify each of the following as an observation, hypothesis, law, or theory.
(a) When wood burns, oxygen is consumed.
(b) Heavier-than-air objects always fall toward the center of Earth.
(c) Matter is composed of atoms.
(d) Crime rates increase when the moon is full.

1.117 Classify each of the following as an observation, an hypothesis, law, or theory.
(a) Bad luck results from walking under a ladder.
(b) Oil floats on water.
(c) Oil floats on water because it is less dense.
(d) Wood burns.

1.118 You observe coins in a fountain and propose the hypothesis that fountains are built with coins in them. Suggest experiments to test this hypothesis.

1.119 You observe a piece of balsa wood floating on water and propose the hypothesis that all wood floats on water because wood is less dense than water. Suggest experiments to test this hypothesis.

Additional Questions

1.120 Rank the following measurements in order from smallest to largest: 1.0×10^{-4} m, 2.0×10^{-5} m, 3.0×10^{-6} km, 4.0×10^{2} mm, 0.0 m, 1.0 m.

1.121 The density of air in a balloon is less at high altitudes than at low altitudes. Explain this difference.

1.122 If the temperature in a room increases from 20.0°C to 30.0°C, what is the temperature change in units of kelvin?

1.123 If you have a sample of zinc and a sample of copper, and both have the same mass, which has the greatest volume?

1.124 Give the symbols for potassium and phosphorus.

1.125 Give the symbols for the following noble gas elements:
(a) helium
(b) neon
(c) argon
(d) krypton
(e) xenon
(f) radon

1.126 The red blood cell (RBC) count for a normal female is about 5 million RBCs per μL (1 μL = 1 mm^3). How many red blood cells are in a gallon of normal female blood?

1.127 Recycling facilities around the world use a variety of techniques to separate different types of recyclable waste. Wet separation is a method that involves processing certain kinds of waste (for example, glass, sand, and metal) based on their densities relative to water. Is this type of separation based on the physical or chemical properties of the recyclable waste?

1.128 These samples of metals have the same mass. Which has the greater density?



A B

1.129 The typical dose of epinephrine at a particular concentration administered to a patient under cardiac arrest is 0.1 mg per kilogram of body weight. If a patient weighs 180 lb, how much epinephrine should be administered?

1.130 About 70 million tons of paper are used per year in the United States. If the U.S. population is about 301.6 million people, how many kilograms of paper are used by the average U.S. citizen?

1.131 During a typical physical exam blood tests to measure cholesterol levels are usually run. A total cholesterol level of 240 mg/dL or above is considered high. What is this threshold level in units of pounds per fluid ounce? If a typical adult has 4 to 6 L of blood, how many pounds of cholesterol are present in the blood of a patient with a total cholesterol level of 260 mg/dL?

1.132 The densities of antifreeze, corn oil, dish detergent, maple syrup, shampoo, and water are 1.13, 0.93, 1.03, 1.32, 1.01, and 1.00, respectively. Which layer is which substance in the figure?



CHAPTER

3

Chemical Compounds

3.1 Ionic and Molecular Compounds

3.2 Monatomic and Polyatomic Ions

3.3 Formulas for Ionic Compounds

3.4 Naming Ionic Compounds

3.5 Naming and Writing Formulas for Molecular Compounds

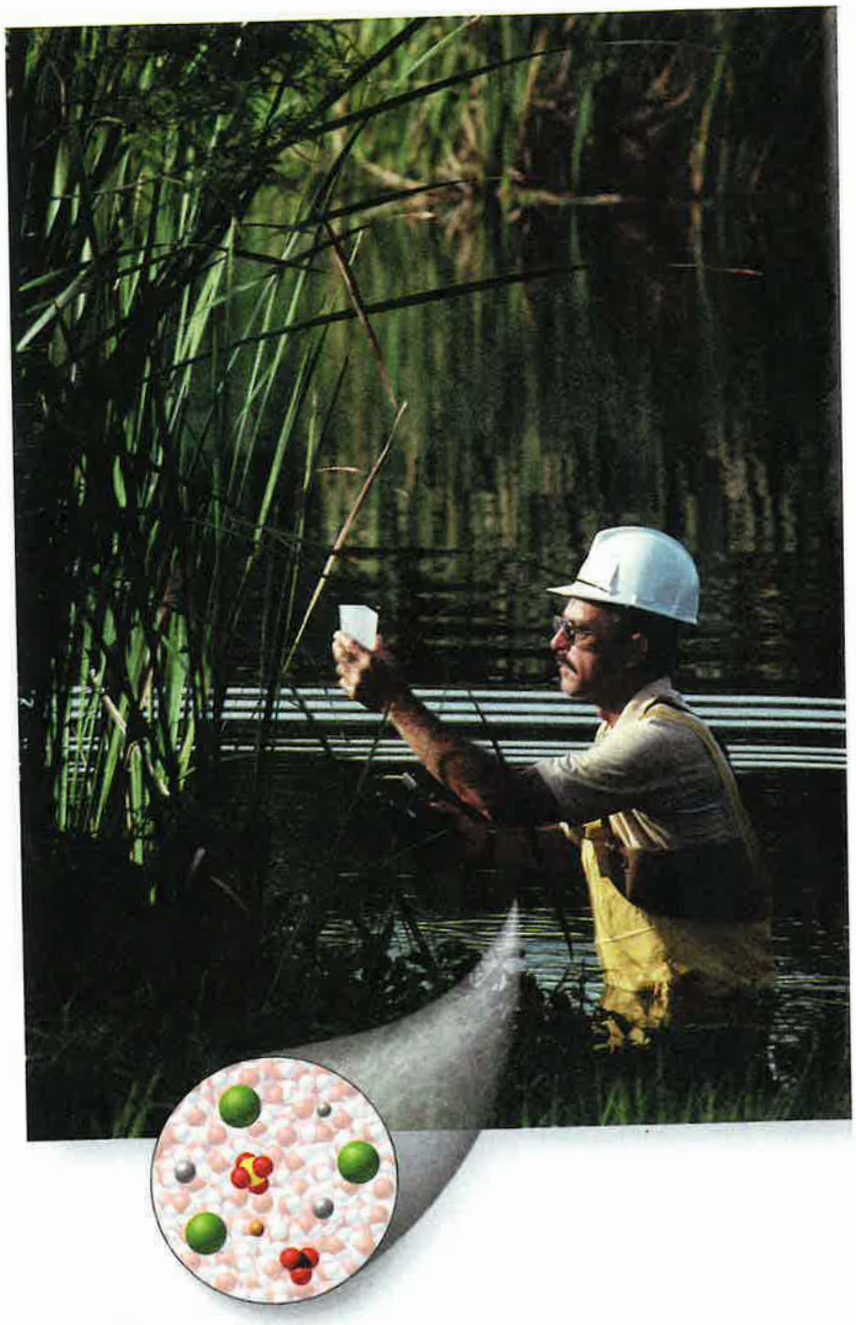
3.6 Acids and Bases

3.7 Predicting Properties and Naming Compounds

Summary

Key Terms

Questions and Problems



Some students taking a chemistry course went on a canoe trip on a local river. They collected water samples from several places along the river for analysis back in their lab. As they paddled downstream, two of the students, Jeff and Megan, began talking about all the different kinds of water they encounter in their daily lives: bottled, mineral, mountain spring, seltzer, tonic, carbonated, lake, ocean, pond, hard, soft, and more. In their precollege science courses, Jeff and Megan recalled learning that water is one of the most important and abundant substances on Earth's surface and that it supports the existence of all living things on this planet. As a liquid, it occurs in the rivers, lakes, and oceans that cover over 70% of Earth's surface. It is also found underground. Groundwater provides about three-fourths of the water used by cities in the United States. Water also exists on Earth in solid form as snow and ice at the polar ice caps and in other cold regions. It is found in the atmosphere as gaseous water vapor and clouds.

The orange juice Jeff and Megan drank as they collected samples along the river is mostly water. The salads they ate for lunch are mostly water, too; and because the day was warm and humid, water vapor condensed on the outside of their iced tea glasses. Before they left for the trip, they showered and brushed their teeth using water. When they rode the bus to the canoe launch site, the vehicle released water into the air from its exhaust pipe. The water formed as a product of combustion in the engine. As they paddled the canoe, Megan and Jeff perspired. "Even our sweat is mostly water," Megan joked.

The human body is about 70% water.

To a chemist, pure water exists as H_2O molecules in which two hydrogen atoms are connected to one oxygen atom (Figure 3.1). In this form, it is colorless, odorless, and tasteless. In nature, water is rarely pure because it has the unique ability to dissolve so many other substances. One of Megan and Jeff's assignments was to analyze their water samples for various materials—suspended and dissolved—that had been carried down the river.

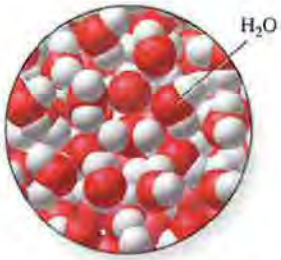


FIGURE 3.1 Pure water, to a chemist, consists of two hydrogen atoms attached to an oxygen atom.

To understand where those materials come from, let's consider how water makes its way from a cloud over Minnesota to the surface of Earth, across the land to the Mississippi River, down the river to New Orleans and the Gulf of Mexico, and ultimately out into the Atlantic Ocean. The water picks up many different substances as it travels, such as minerals and organic matter that will later be absorbed or ingested by plants and animals. Falling through the air as rain, it dissolves gases from the atmosphere. As it percolates through the ground and runs down streambeds, it dissolves a variety of minerals and gases. As it flows in smaller rivers that feed the Mississippi, it gathers more minerals. Finally it becomes "salty" as it mixes with seawater in the Gulf of Mexico. Water from the surface enters the atmosphere as puddles and ponds evaporate. Evaporation from rivers, lakes, and oceans also contributes to atmospheric water. Transpiration (water loss) from plants adds water to the atmosphere, too. In the air, gaseous water condenses to form clouds, beginning the process all over again (Figure 3.2).

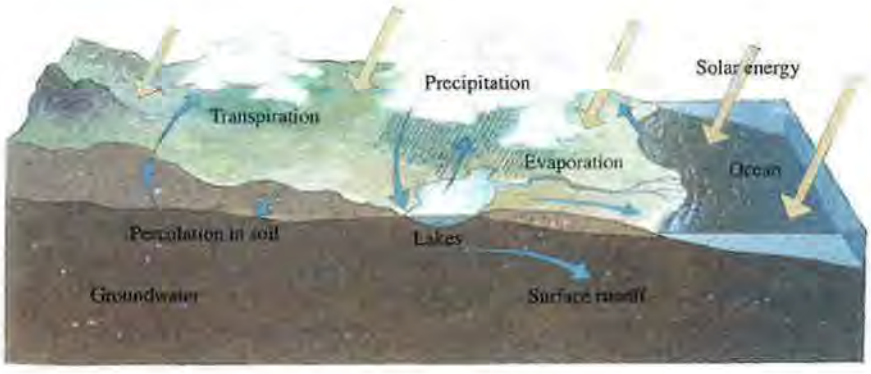


FIGURE 3.2 Water is transported from place to place on Earth in a series of processes called the water cycle. As water flows across the land or falls through the air, many substances end up suspended or dissolved in it.

How does *dissolving* differ from *melting*?

From rain cloud to ocean, water is always in the form of a solution, with water acting as the solvent. Numerous substances dissolved in the water provide nutrients to plants and animals. For example, many of the minerals needed by living organisms are present in water as ions. They include ions of iron, calcium, potassium, and magnesium. The ability of water to dissolve other substances is very important. Many of the topics we discuss throughout this book involve the chemical behavior of substances dissolved in water.

Many of the substances that dissolve in water can be classified into categories of chemical compounds. The categories vary greatly in their physical and chemical properties. In this chapter we'll examine the characteristics and behavior of these compounds and look at different ways of naming them.

Questions for Consideration

3.1 How do ionic compounds differ from molecular compounds?
3.2 What kinds of ions are in ionic compounds?
3.3 What do formulas for ionic compounds represent?
3.4 How are ionic compounds named?
3.5 What do formulas for molecular compounds represent and how are they named?
3.6 What are some common acids and bases and how are they named?
3.7 How do names of compounds communicate their classification and properties?

3.1 | IONIC AND MOLECULAR COMPOUNDS

When they returned to the lab, Jeff and Megan performed several tests on their samples, including one to determine the water's electrical conductivity. They used a simple device like that shown in Figure 3.3 to see if substances dissolved in the river water allow a complete electric circuit to operate. Notice in Figure 3.3B the snipped wire that leads down from the lightbulb. If the snipped wires were to touch, the circuit would be complete and the bulb would glow. Another way of completing the circuit is to immerse the wires in a solution that conducts electricity. For an electric current to pass through a solution, ions must be present.



A Circuit complete **B** Circuit not complete **C** Circuit complete

FIGURE 3.3 (A) When electricity flows around a complete circuit, the bulb lights. (B) Snip any wire and the circuit is broken. (C) When the ends of the snipped wire are placed in a solution of a substance that contains ions, the bulb glows again. This is a simple procedure for testing the conductivity of a substance dissolved in water.



A Pure water B Aqueous sugar solution C Aqueous NaCl solution

FIGURE 3.4 When electrical conductivity is tested with a lightbulb apparatus as shown in Figure 3.3, (A) pure water and (B) a sugar solution do not appear to complete the circuit, but (C) a sodium chloride solution does—the bulb glows.

Megan and Jeff determined that the water conducted an electric current. To investigate further, they tested several different materials in their lab. Consider the three images shown in Figure 3.4. When the electrical conductivity of pure water is tested, the lightbulb does not glow. Pure water is not a good enough conductor to complete the circuit, and neither is a solution of sugar in water. However, when electricity is passed through a sodium chloride, NaCl, solution, the bulb glows brightly. For electric current to flow through an aqueous solution, ions must be present. NaCl provides them, while pure water and sugar water do not—at least not in sufficient quantities to carry the electric current. A substance like NaCl that releases ions when dissolved in water is an **electrolyte** (Figure 3.5). One that does not, such as sugar, is a **nonelectrolyte**. Electrolyte solutions conduct electricity, while nonelectrolyte solutions do not.

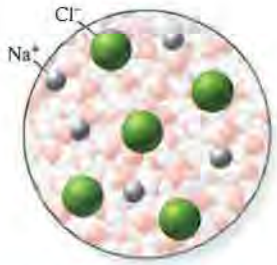


FIGURE 3.5 Sodium chloride, NaCl, is an electrolyte because it separates into ions when dissolved in water.

We can be more specific by distinguishing among strong electrolytes, weak electrolytes, and nonelectrolytes. Consider the solutions shown in Figure 3.6. Which compounds appear to break apart as ions in solution? Do some of the compounds produce more ions than others? The process by which a compound dissolves in water to produce ions is called *dissociation* or *ionization* depending on the type of compound dissolving. Because a **strong electrolyte** dissociates extensively in water, it conducts electricity well. Notice in Figure 3.6 that sodium chloride, NaCl, and hydrochloric acid, HCl, dissociate extensively. They are good conductors of electricity. Other substances, such as acetic acid, CH_3CO_2H, dissociate only partially in water. Such a substance does not conduct electricity well and is called a **weak electrolyte**. The solution containing methanol, CH_3OH, does not conduct electricity. Methanol does not dissociate into ions at all. It is a nonelectrolyte.

ANIMATION: Strong Electrolyte, Weak Electrolyte, and Nonelectrolyte

Many experiments, including those conducted by Jeff and Megan, have shown that, in general, compounds can be classified into two categories. An **ionic compound** (sometimes generically referred to as a *salt*) consists of oppositely charged cations and anions in proportions that give electrical neutrality. Ionic compounds are easily identified because they usually consist of ions from a *metal* with ions from a *nonmetal*. Sodium chloride, which is table salt, and is composed of Na^+ and Cl^- ions, is an example of an ionic compound (Figure 3.7A). A **molecular compound**, on the

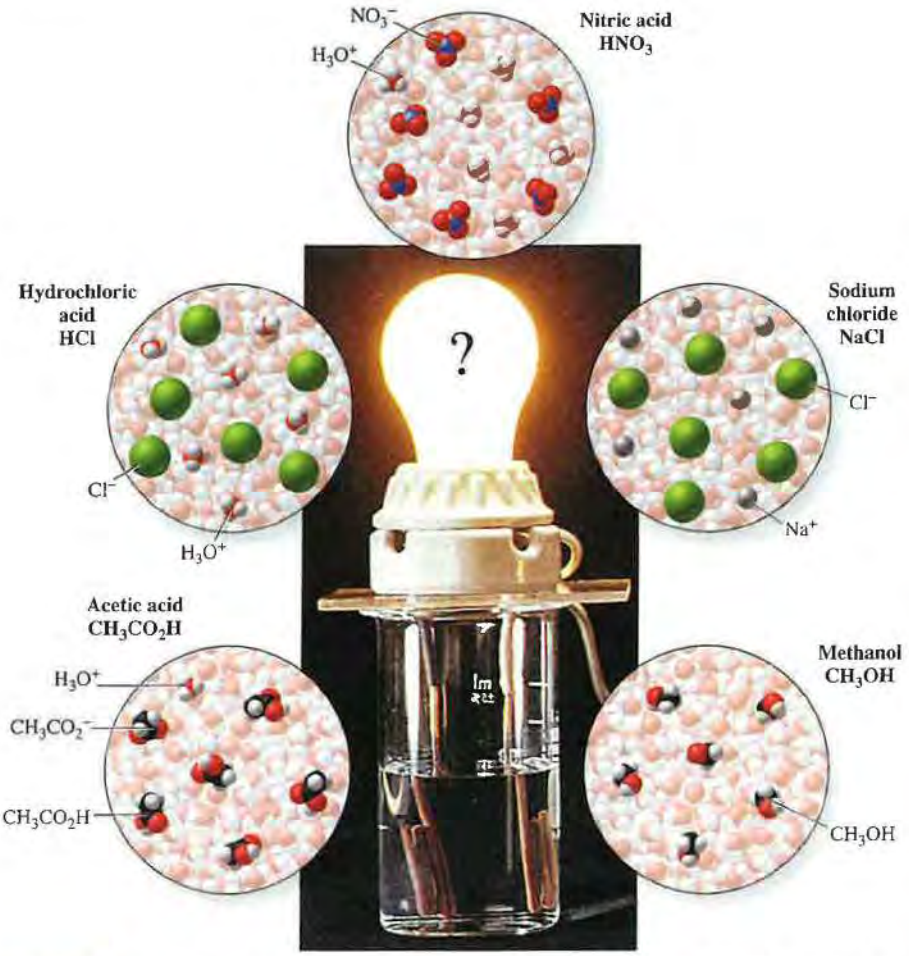
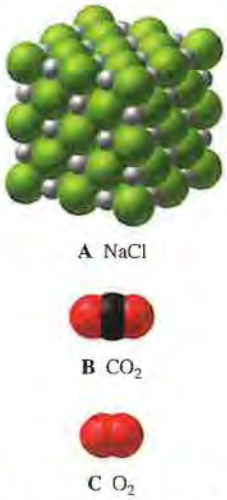


FIGURE 3.6 For each of the solutions, inspect the molecular-level images in which each type of particle is labeled. If any of these substances dissociate into ions, separate particles of single atoms (or atom groups) are visible. Which of the compounds dissociate completely in solution? Which only partially dissociate? Which do not dissociate at all? The acids do not simply dissociate, but form H_3O^+ ions in solution.

other hand, is composed of atoms from two or more *nonmetals*. It exists as a discrete unit of atoms held together in a molecule. Carbon dioxide, CO_2, is a molecular compound (Figure 3.7B), and O_2 is a molecular element (Figure 3.7C).



A NaCl

B CO_2

C O_2

FIGURE 3.7 Sodium chloride, NaCl, is an ionic compound. Carbon dioxide, CO_2, is a molecular compound. Elemental oxygen, O_2, is a molecular element.

EXAMPLE 3.1 | Ionic and Molecular Compounds

Based on their formulas, which of the following compounds are ionic?

(a) KCl (b) CO_2 (c) CaO (d) CCl_4

Solution:

We can determine if a compound is ionic by looking at the elements that compose it. An ionic compound is usually composed of ions from a metal and a nonmetal. Two of the compounds meet this criterion, KCl and CaO.

Practice Problem 3.1

Which of the compounds listed in the example are molecular?

Further Practice: Questions 3.7 and 3.8 at the end of the chapter

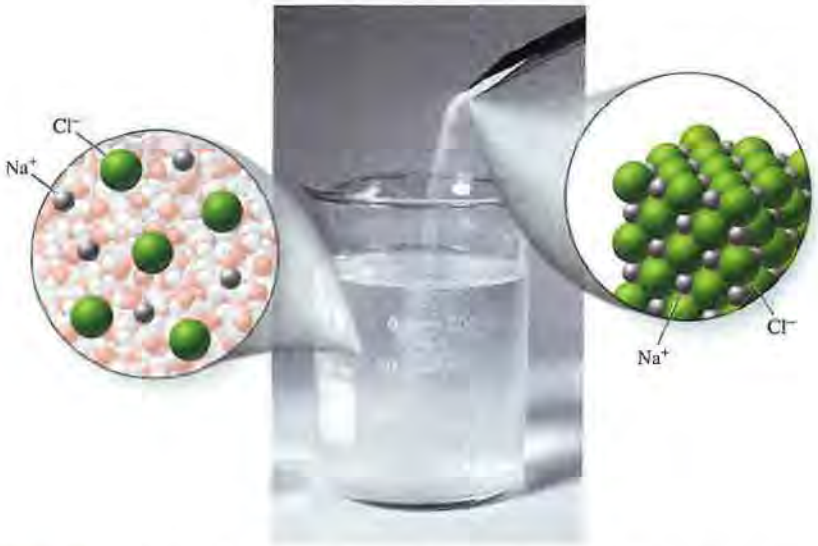


FIGURE 3.8 In solid NaCl, the ions are arranged in a crystal lattice. Once sodium chloride dissolves in water, the ions of sodium and chlorine separate and become surrounded by water molecules.

To see how ionic and molecular compounds differ, let's first consider their appearance and what happens to them in aqueous solution. Ionic compounds that dissolve in water separate into the independent ions that make them up. When a compound composed of ions separates into its component ions in water, the process is described as *dissociation*. In contrast, molecular elements and compounds generally retain their molecular structure. The atoms contained in the molecule usually stay together when the substance dissolves in water.

ANIMATION: Sodium Chloride Dissolving in Water

Notice in Figure 3.8 the ordered arrangement of cations and anions in the ionic solid sodium chloride. This ordered arrangement is one type of solid structure called a *crystal structure* or *crystal lattice*. The strong attraction between the oppositely charged ions in an ionic compound—along with the pattern in which they arrange themselves—gives ionic solids their unique properties. They are solids at room temperature, have very high melting points, and are brittle.

The crystal lattice of ionic solids breaks down when they dissolve in water. The solid dissociates into ions. Compare the molecular-level image of solid sodium chloride in Figure 3.8 to the representation of it dissolved in water. How do the ions of sodium and chlorine differ in the pictures?

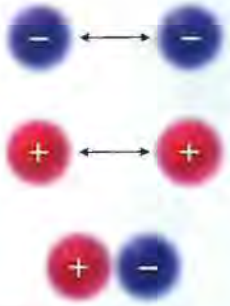
We represent the process of sodium chloride dissolving with the equation

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

where the H_2O written above the arrow indicates that water participates in the process but is not permanently changed. The (aq) after Na^+ and Cl^- indicates that these ions are in aqueous solution.

In contrast to ionic compounds, most molecular compounds remain as molecules when dissolved in water. When not in solution, molecular compounds can be gas, liquid, or solid at room temperature. They have much lower melting and boiling points than ionic compounds. Carbon dioxide, CO_2, is a molecular compound. Because carbon dioxide is composed of molecules of uncharged atoms, not ions, its appearance and behavior are different from those of an ionic compound (Figure 3.9). Molecular elements also differ from ionic compounds. For example, elemental oxygen is diatomic; that is, it exists as two oxygen atoms bound together in a molecule, represented as O_2. When it dissolves in water, the oxygen atoms remain together as a molecule (Figure 3.10).

Like charges repel. Opposite charges attract.



Ozone is actually a triatomic oxygen molecule, O_3. The ozone layer within the stratosphere (the atmospheric layer that extends from 10 to 50 km above Earth) helps protect living things on the surface by absorbing much of the harmful ultraviolet radiation in sunlight.



FIGURE 3.9 While molecular compounds like carbon dioxide are composed of molecules, not ions, some can also form solids. At room temperature, CO_2 readily goes from the solid phase directly to a gaseous state. (Other gas particles are not shown in the molecular-level image of the gaseous state.)



FIGURE 3.10 Many molecular substances such as oxygen, O_2, remain as intact molecules when dissolved in water.

Other differences between ionic and molecular compounds are summarized in Table 3.1. To explain these differences, we can compare the molecular-level structures of ionic and molecular compounds. For example, compare solid NaCl to solid NH_3 in Figure 3.11 and imagine what is involved when these two compounds melt. The process of melting for any substance requires the dispersion of the components of the solid into a more random arrangement. In the case of NaCl, the components are cations and anions, and its melting point is 801°C. In the case of NH_3, the components are individual NH_3 molecules, and its melting point is only –77.7°C. The greater heat energy required to melt the ionic compound indicates that the force of attraction between oppositely charged ions is much stronger than the attractive forces between molecules when they are in the solid state.



FIGURE 3.11 Strong attractive forces between oppositely charged cations and anions must be overcome for solid NaCl to melt. Molecules of NH_3 in the solid state are not held together by oppositely charged particles. Instead, weak forces between the NH_3 molecules hold them together in the crystal.

TABLE 3.1 | Properties of Ionic and Molecular Compounds Compared

Ionic Compounds	Molecular Compounds
crystalline solid	gas, liquid, or solid
hard, brittle solid	soft solid
very high melting point	low melting point
very high boiling point	low boiling point
high density	low density
strong electrolyte in aqueous solution	weak electrolyte or nonelectrolyte in aqueous solution
electrical conductivity good when compound is molten	electrical conductivity poor in pure form

EXAMPLE 3.2 | Properties of Ionic and Molecular Compounds

Which would you expect to have the lower melting point, KCl or CCl_4?

Solution:

Ionic compounds consist of oppositely charged ions that have a strong force of attraction between them. These compounds require significantly higher temperatures for melting. It would take a lot more energy to break apart the ions of KCl than to separate CCl_4 molecules from each other, so CCl_4 should have the lower melting point.

Practice Problem 3.2

Which would you expect to have the higher boiling point, KCl or CCl_4?

Further Practice: Questions 3.9 and 3.10 at the end of the chapter

In Example 3.2, the compounds listed are composed of only two elements. They are binary compounds. A **binary compound** is a compound containing atoms or ions of only two elements. Many chemical compounds are binary. Some ionic compounds contain atoms of two different nonmetals that form a single, binary ion. You'll learn more about these special ions in Section 3.2.

Binary compounds are composed of only two elements, but often are composed of more than two atoms (e.g., N_2O_5).

3.2 | MONATOMIC AND POLYATOMIC IONS

When Jeff and Megan tested their samples from the river, they discovered a lot of sodium, magnesium, and iron in the water. These elements were not present in the form of solid metals. They were dissolved in the water as cations. In addition, Megan and Jeff found several anions. When they completely evaporated all the water from a sample by boiling it off, a variety of compounds remained as a solid residue in the flask. Most of them were ionic compounds.

Monatomic Ions

To better understand the composition of ionic compounds, let's look more closely at the kinds of ions that make them up. In Chapter 2 we described how the charges on the ions of some main-group elements can be predicted from their positions in the periodic table. For example, find magnesium and sulfur in the periodic table. Based on their positions, what charge would you expect for their ions? A magnesium atom has two more electrons than an atom of its nearest noble gas, neon. A magnesium ion will have the same number of electrons as a neon atom, which is two less than the magnesium atom. We represent that ion as Mg^{2+}. In contrast, a sulfur atom contains two fewer electrons than the nearest noble gas, argon. A sulfur ion will have the same number of electrons as an argon atom, which is two more than a sulfur atom. A sulfur atom becomes negatively charged with two extra electrons, represented as S^{2-}. The ions of magnesium and sulfur are monatomic. A **monatomic ion** is an ion of a single atom.

Several common monatomic ions are shown in Figure 3.12. Many ions of the main-group elements, those on the far left or far right of the periodic table, can be predicted from their positions. However, the charges of ions for elements in the middle, especially the transition metals, cannot be predicted from their group numbers. These elements often form more than one kind of ion.

Nomenclature is a system of naming. To distinguish one chemical entity from another, unique names are assigned to all elements, ions, and compounds. Just as

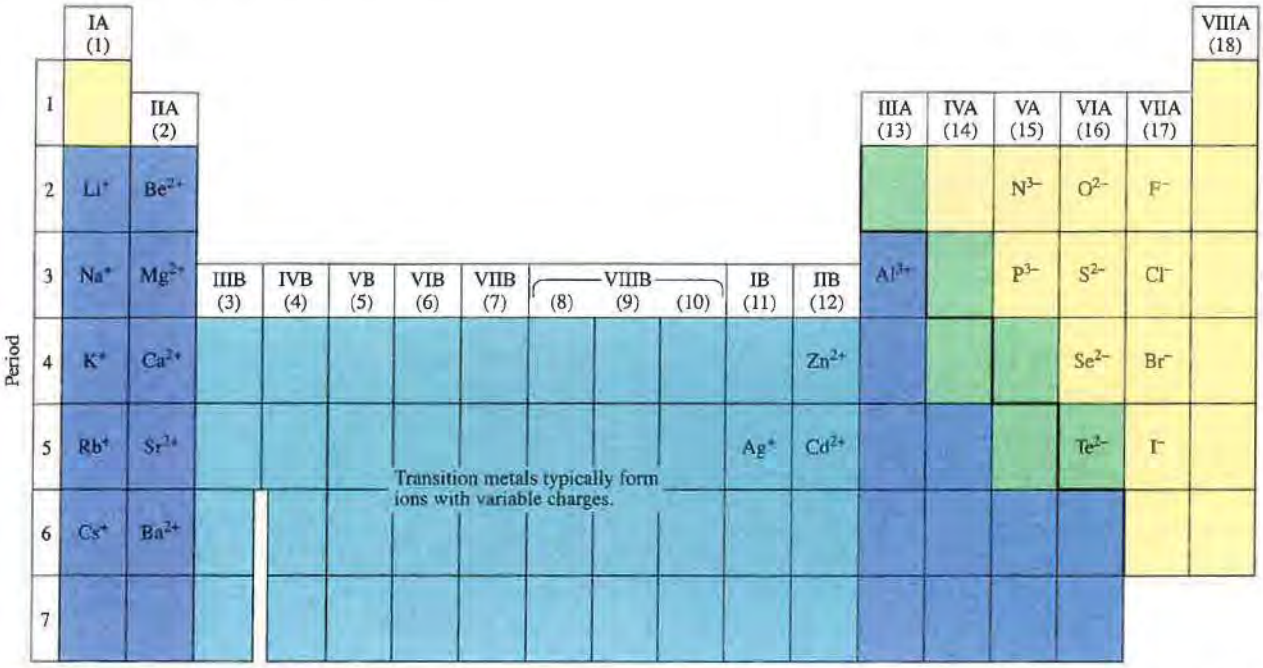


FIGURE 3.12 The charges on the common monatomic ions of many main-group elements correlate with the elements' positions in the periodic table. Charges on the other elements are not easily predicted from their positions. A few of the transition metals exhibit only one charge.

TABLE 3.2 | Common Monatomic Ions

Formula of Ion	Name of Ion	Formula of Ion	Name of Ion	Formula of Ion	Name of Ion
O^{2-}	oxide ion	Al^{3+}	aluminum ion	K^+	potassium ion
Na^+	sodium ion	Mg^{2+}	magnesium ion	F^-	fluoride ion
N^{3-}	nitride ion	Cl^-	chloride ion		

elements have unique names and symbols, ionic compounds and ions have their own names—and a special set of rules for correct naming. Several common monatomic ions are listed in Table 3.2. Try to determine some of the rules that were used to name them.

Did you figure out how to name ions? Did you determine when the *–ide* ending is used and how to use it? Now try to apply your rules in Example 3.3.

EXAMPLE 3.3 | Formulas and Names of Ions

Based on their positions in the periodic table, predict the charge, write the formula, and name the ions that the following elements are expected to form.

(a) lithium
(b) sulfur

Solution:

(a) An atom of lithium is expected to lose one electron to have the same number as helium. Its ion is represented with the formula Li^+ and is named *lithium ion.*
(b) Sulfur is expected to gain two electrons to have the same number as argon. The representation for the ion of sulfur is S^{2-} and is named *sulfide ion.*

Practice Problem 3.3

Based on their positions in the periodic table, predict the charge, write the formula, and name the ions that the following elements are expected to form.

(a) barium
(b) bromine

Further Practice: Questions 3.11 and 3.12 at the end of the chapter

If you had trouble establishing rules for naming ions, the following summary may help. Monatomic anions are named as the first part of the element name (called the root of the name) and *–ide* is added as a suffix. For example, S^{2-} is known as sulf*ide* ion. Monatomic cations are not given special endings to their root names, so we refer to them as an ion of their element name; for example, Na^+ is known as *sodium ion.*

Polyatomic Ions

Not all ions are monatomic. Some are composed of multiple atoms. A **polyatomic ion** is an ion containing two or more atoms, usually of more than one element. An example is the nitrate ion, NO_3^-. Nitrate ions are discrete units of three oxygen atoms and one nitrogen atom, as represented in Figure 3.13. Although the formula for nitrate ion might suggest that the oxygen atoms are attached to each other, they are actually attached to a central nitrogen atom. This polyatomic ion bears a charge of 1– because it has one more electron than the total number provided by the nitrogen and oxygen atoms.

The most common polyatomic ions are anions that contain oxygen attached to some other element. Such an anion is called an **oxoanion**. Oxoanions are typically combinations of oxygen with a nonmetal, although some contain metals. For a given nonmetal, there are often two or more common oxoanions. When this is the case, the two oxoanions differ in the number of oxygen atoms surrounding the nonmetal atom. For example, there are two known oxoanions of sulfur, both with a 2– charge. One contains a central sulfur atom with three oxygen atoms attached to it. The other has a central sulfur atom with four oxygen atoms. These oxoanions are represented in Figure 3.14 with the formulas SO_3^{2-} (sulf*ite* ion) and SO_4^{2-} (sulf*ate* ion).

Some water supplies contain a variety of aqueous ions, including sulfate, carbonate, phosphate, and chloride ions, as well as ions of calcium, magnesium, iron, and aluminum. Water containing the cations Ca^{2+}, Mg^{2+}, and sometimes Fe^{3+}, along with anions such as Cl^-, SO_4^{2-}, and HCO_3^-, is called *hard water.* When Megan and Jeff boiled away their water samples, they noticed a residue left on the glassware. They tested it and found that carbonate ions, CO_3^{2-}, were present. The hard water deposits were mostly metal carbonates, $CaCO_3$ and $MgCO_3$. They had noticed similar residues at home on cooking pans and drinking glasses from which water had boiled or evaporated. Carbonate deposits also build up in pipes (Figure 3.15). Although their formulas are more complex than those of binary compounds, $CaCO_3$ and $MgCO_3$ behave like binary ionic compounds. When $CaCO_3$ dissolves in water, Ca^{2+} and CO_3^{2-} ions form (Figure 3.16). When the water is removed by evaporation or boiling, metal carbonates in solid form stay behind.

Nearly all polyatomic ions are negatively charged. An exception is the common polyatomic cation NH_4^+ (called ammonium ion). Like monatomic ions, polyatomic ions have their own rules for naming. Several common ones are shown in Table 3.3. Can you figure out the rules used to name them?

Now apply your rules to naming the polyatomic ions in Example 3.4.



NO_3^-

FIGURE 3.13 In the nitrate ion, three oxygen atoms surround one nitrogen atom.



S^{2-}
Sulf*ide*



SO_3^{2-}
Sulf*ite*



SO_4^{2-}
Sulf*ate*

FIGURE 3.14 Anions containing sulfur differ in the number of oxygen atoms attached to the sulfur. All have a 2– charge.

In a formula for a binary molecular compound or polyatomic ion, the element present in the least amount is usually central to the atoms of the other element. The central atom is usually listed first in the formula.

Many commercial and domestic water-softening products and devices are available for removing the Ca^{2+}, Mg^{2+}, and Fe^{3+} ions in hard water.



FIGURE 3.15 Over time, solid buildup of carbonates can close off pipes.

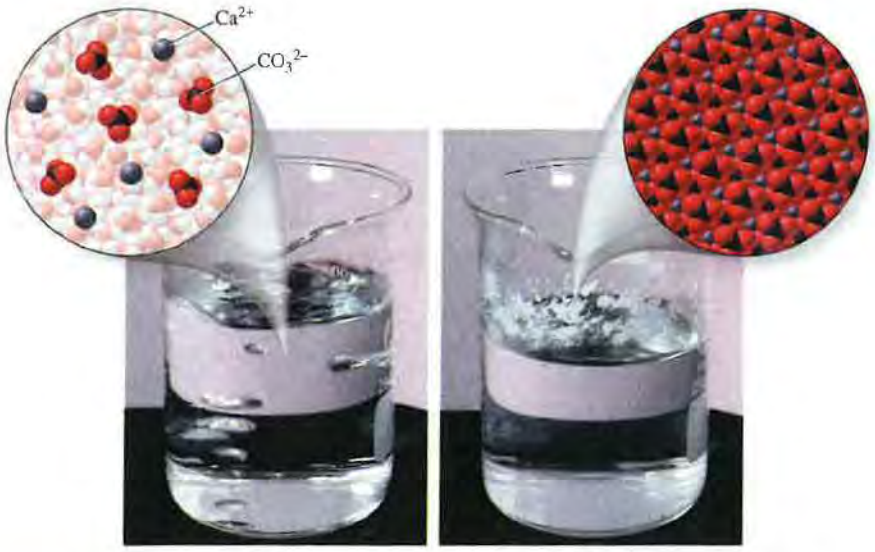


FIGURE 3.16 A solution of $CaCO_3$ contains a small amount of Ca^{2+} and CO_3^{2-} ions. Boiling the water from the solution leaves behind solid deposits of calcium carbonate.

TABLE 3.3 | Common Oxoanions

Formula of Ion	Name of Ion
NO_3^-	nitrate ion
NO_2^-	nitrite ion
SO_4^{2-}	sulfate ion
SO_3^{2-}	sulfite ion
ClO_3^-	chlorate ion
ClO_2^-	chlorite ion

EXAMPLE 3.4 | Formulas and Names of Polyatomic Ions

There are two oxoanions of arsenic. One has three oxygen atoms attached to an arsenic atom and has a 3− charge. The other has the same charge and four oxygen atoms attached to an arsenic atom. Write the formula for the arsenite ion.

Solution:

From Table 3.3 we can tell that for a particular nonmetal oxoanion, the one that contains fewer oxygen atoms gets an *–ite* ending. We represent the arsenite ion as AsO_3^{3-}.

Practice Problem 3.4

Write the formula for the arsenate ion.

Further Practice: Questions 3.23 and 3.24 at the end of the chapter

If you had trouble establishing rules for naming oxoanions, the following summary may help. If there are only two oxoanions of an element, the one with the greater number of oxygen atoms is named by combining the root of the nonmetal element name and the ending *–ate*. The oxoanion with the lesser number of oxygen atoms uses the ending *–ite*. Thus, sulf*ate* is the sulfur oxoanion with the formula SO_4^{2-} and sulf*ite* is the name of SO_3^{2-}. Formulas and charges for several of the most common oxoanions are shown in Figure 3.17. Do you see any patterns in the ions shown?

Some nonmetals form more than two oxoanions. For example, chlorine forms four known oxoanions. They are named as before with the *–ate* or *–ite* suffixes. However, the prefix *per–* is placed before the root of the element name and *–ate* is added to the end for the oxoanion that contains the most oxygens. The prefix *hypo–* and the ending *–ite* are used for the anion with the fewest oxygens. These, plus other polyatomic ions whose names do not fit regular patterns, are listed in Table 3.4.

Too much nitrogen and phosphorus, in the form of nitrates and phosphates, in water—from fertilizers, sewage, or livestock runoff—stimulates the growth of algae. These algal growth spurts, or blooms, soon die off. Bacteria decompose the dead plants, consuming oxygen in the process. The reduction of dissolved oxygen in the water kills fish and other aquatic organisms.

B BO_3^{3-} borate	**C** CO_3^{2-} carbonate	**N** N^{3-} nitride NO_3^- nitrate NO_2^- nitrite	**O** O^{2-} oxide O_2^{2-} peroxide	**F** F^- fluoride No oxoanions
	Si SiO_4^{4-} silicate	**P** P^{3-} phosphide PO_4^{3-} phosphate	**S** S^{2-} sulfide SO_4^{2-} sulfate SO_3^{2-} sulfite	**Cl** Cl^- chloride ClO_4^- perchlorate ClO_3^- chlorate ClO_2^- chlorite ClO^- hypochlorite
		As As^{3-} arsenide AsO_4^{3-} arsenate AsO_3^{3-} arsenite	**Se** Se^{2-} selenide SeO_4^{2-} selenate SeO_3^{2-} selenite	**Br** Br^- bromide BrO_4^- perbromate BrO_3^- bromate BrO_2^- bromite BrO^- hypobromite
			Te TeO_4^{2-} tellurate TeO_3^{2-} tellurite	**I** I^- iodide IO_4^- periodate IO_3^- iodate IO_2^- iodite IO^- hypoiodite

FIGURE 3.17 Formulas and charges of some monatomic anions and oxoanions. Do you see any patterns?

TABLE 3.4 | Important Polyatomic Ions

1– Ions		2– Ions	
nitrate	NO_3^-	chromate	CrO_4^{2-}
nitrite	NO_2^-	dichromate	$Cr_2O_7^{2-}$
bicarbonate (hydrogen carbonate)	HCO_3^-	sulfate	SO_4^{2-}
perchlorate	ClO_4^-	sulfite	SO_3^{2-}
chlorate	ClO_3^-	carbonate	CO_3^{2-}
chlorite	ClO_2^-	oxalate	$C_2O_4^{2-}$
hypochlorite	ClO^-	peroxide	O_2^{2-}
cyanide	CN^-	**3– Ions**	
hydroxide	OH^-	phosphate	PO_4^{3-}
acetate	$CH_3CO_2^-$	borate	BO_3^{3-}
permanganate	MnO_4^-	**1+ Ion**	
		ammonium	NH_4^+

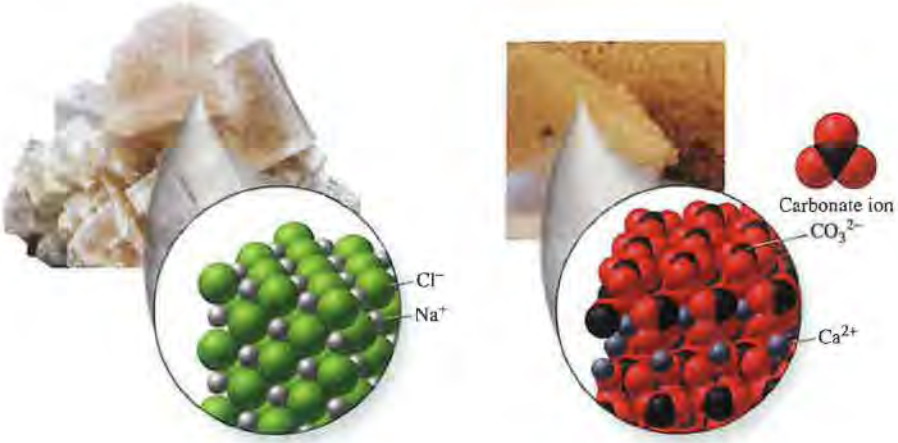


FIGURE 3.18 Ionic compounds containing polyatomic ions such as calcium carbonate, $CaCO_3$, exist in a regular array of cations and anions. Notice the similarities between sodium chloride and calcium carbonate.

In their solid form, ionic compounds exist in an organized array of alternating positive and negative ions. Calcium carbonate, $CaCO_3$, the hard water residue that is also a major component of limestone, is another example of an ionic compound. Although its formula might suggest that it is nothing like sodium chloride, it does have similar properties. It is composed of positive and negative ions: the calcium ion, Ca^{2+}, and the carbonate ion, CO_3^{2-}. Notice in Figure 3.18 that calcium carbonate forms an ordered array of alternating cations and anions similar to sodium chloride's crystal lattice structure. Instead of a monatomic anion, the anion in this case is CO_3^{2-}. As indicated by the molecular-level representation, a polyatomic ion behaves as a unit.

3.3 | FORMULAS FOR IONIC COMPOUNDS

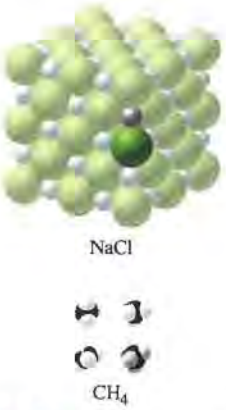


FIGURE 3.19 Ionic compounds such as NaCl do not exist as discrete molecules. The highlighted region of the NaCl structure is the formula unit for this compound. Molecular compounds such as CH_4 do exist as a discrete unit.

Some of the ionic compounds that Megan and Jeff isolated from their river water sample were $MgCl_2$, $NaNO_3$, $CaSO_4$, and $FeCl_3$. What do these formulas represent? How are they determined? Whether in a solid or in solution, the ions that make up an ionic compound do not exist as molecules. Contrast sodium chloride and methane in Figure 3.19. The lattice structure of sodium and chloride ions is different from the molecular structure of CH_4. Methane exists as discrete molecules, each composed of one carbon atom and four hydrogen atoms.

If ionic compounds don't exist as molecules, then how can we write formulas for them? Combinations of oppositely charged ions, either in solid form or in solution, are present in quantities that result in a net charge of zero. That is, the number of positive charges equals the number of negative charges. Formulas for ionic compounds are written to reflect this fact. The sum of the positive charges must equal the sum of the negative charges:

Total positive charge from cations	+	total negative charge from anions	=	zero net charge

Based on their positions in the periodic table, we can predict that the charge on sodium in an ionic compound is 1+ and the charge on chlorine in an ionic compound is 1−. For the compound to be electrically neutral, the two ions have to be present in a one-to-one ratio, or Na_1Cl_1. When there is only one of a given atom or ion in a formula, the subscript 1 is understood, so the formula for sodium chloride can be written simply as NaCl.

Compare this formula to the highlighted region of the sodium chloride structure in Figure 3.19. It has repeating units of one sodium ion and one chloride ion. Because an ionic compound does not exist as a molecule, its formula represents its smallest repeating unit. We call this repeating unit a **formula unit.** The formula unit of sodium chloride is NaCl. How do you know to put the sodium first? Generally, chemical names start with the name of the element that is farther to the left or farther down on the periodic table. Most often, the name or symbol of a metal is written first.

ANIMATION: Sodium Chloride

For ionic compounds, the formula represents the *ratio* of ions of each kind, not an exact count of ions. Consider another example, sodium sulfide. Sodium ion should be Na^+ and sulfide ion should be S^{2-} according to their positions in the periodic table. Two sodium ions are required to provide the 2+ charge needed to balance the 2− charge of the sulfide ion, so these ions will combine to form a compound with the formula Na_2S.

Formulas for ionic compounds containing polyatomic ions must also be written so that the sum of the positive charges from the cations equals the sum of the negative charges from the anions. For example, consider calcium carbonate (Figure 3.20). From the periodic table we know that the calcium ion has a charge of 2+ (Ca^{2+}). From Table 3.4 we know carbonate ion has a charge of 2− (CO_3^{2-}). Therefore, the formula unit is written $CaCO_3$. The calcium appears first in both the name and the formula because it is the cation. We cannot deduce the charge on a polyatomic ion from the periodic table. We must remember it or look it up.

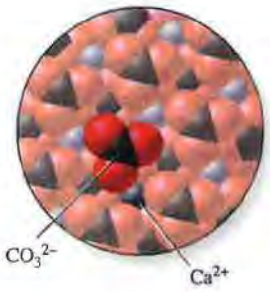


FIGURE 3.20 The formula unit of calcium carbonate is $CaCO_3$.

As another example, consider a compound containing calcium ion and nitrate ion. Based on its position in the periodic table we expect the calcium ion to carry a charge of 2+ (Ca^{2+}). The nitrate ion is known to have a charge of 1− (NO_3^-). Two nitrate ions are required for every calcium ion to write a formula for the compound that is neutral: $Ca(NO_3)_2$. The parentheses around the formula for the polyatomic ion show that the subscript applies to the whole unit. If only one polyatomic ion is represented in a formula unit, the parentheses are not needed. For examples, see Figure 3.21.

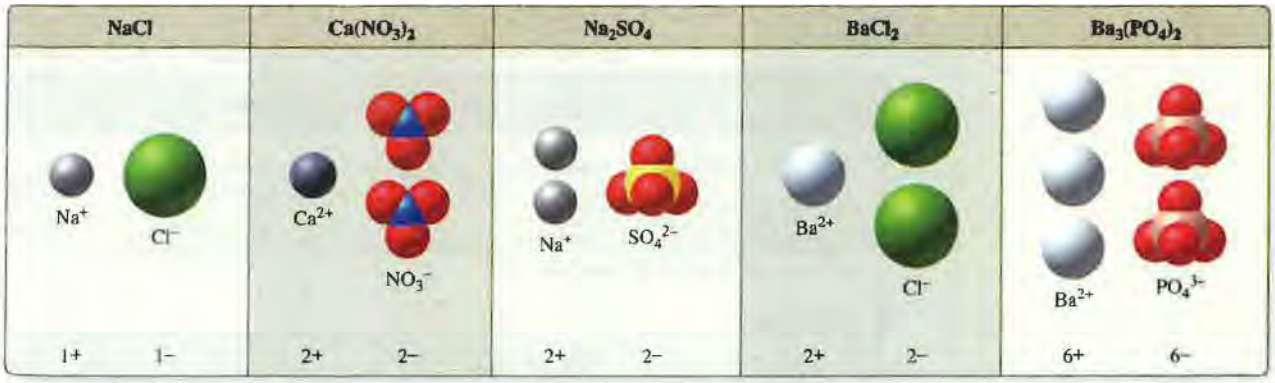
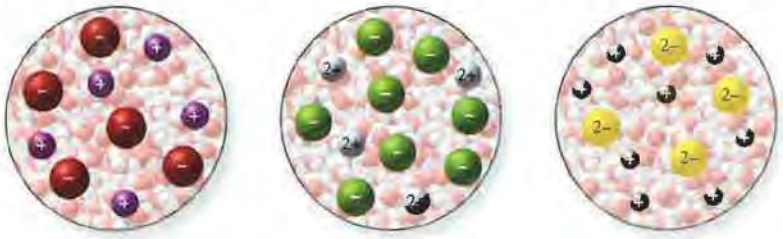


FIGURE 3.21 Ionic compounds are formed by combinations of cations and anions in proportions that lead to electrical neutrality. The images represent a formula unit of each compound.

EXAMPLE 3.5 | Ionic Compounds

The given images represent aqueous solutions of ionic compounds. Match the images to the following formulas: $BaCl_2(aq)$, $Na_2S(aq)$, and $KBr(aq)$.



A B C

Solution:

Based on their positions in the periodic table, barium is expected to have a charge of 2+ and the chloride ion is expected to be 1−. Image B corresponds to $BaCl_2$. Based on its position in the periodic table, we expect sodium ion to have a charge of 1+. The sulfide ion has a charge of 2−. Image C corresponds to Na_2S. Based on their positions in the periodic table, the potassium ion is expected to have a charge of 1+ and the bromide ion is expected to be 1−. Image A corresponds to KBr. Notice that in all of the images ions are present in appropriate ratios of cations and ions.

Practice Problem 3.5

Suppose an ionic compound containing magnesium and chloride ions is dissolved in water. Beginning with a representation for five magnesium ions, draw a picture of a solution containing magnesium and chloride ions that is electrically neutral. You can omit the water molecules.

Further Practice: Questions 3.27 and 3.28 at the end of the chapter

In Example 3.5 we matched formulas with molecular-level representations. We can also write formulas from ion names. First, we determine if a polyatomic ion is present. Next, we determine the charge on the ions that make up the compound. If the ion is monatomic, the charge can often be predicted from the periodic table. If a polyatomic ion is present, we must recall its charge. Once we know the charges on the cations and anions in the compound, we can write a correct formula, as Example 3.6 demonstrates.

EXAMPLE 3.6 | Formulas for Ionic Compounds

Write the formulas for compounds containing the following ions:

(a) calcium ion and nitride ion
(b) barium ion and nitrate ion
(c) potassium ion and sulfate ion

Solution:

(a) Calcium ion and nitride ion are monatomic. Their charges can be determined from their positions in the periodic table. Calcium ion is expected to have a 2+ charge, and nitride ion is expected to have a 3− charge. Combining three calcium ions with two nitride ions yields a formula that has equal amounts of positive and negative charges: Ca_3N_2.
(b) Barium ion is a monatomic cation. Nitrate ion is a polyatomic anion. We can use the periodic table to predict the charge of the cation. The barium ion

is expected to have a 2+ charge. From Table 3.4 we know that nitrate has a charge of 1−. It takes two nitrate ions for every one barium ion to give equal amounts of positive and negative charges: $Ba(NO_3)_2$.

(c) Potassium ion is a monatomic cation. Sulfate ion is a polyatomic anion. Based on its position in the periodic table, potassium ion is expected to have a 1+ charge. From Table 3.4 we know that the charge on sulfate ion is 2−. It takes two potassium ions for every one sulfate ion to yield equal amounts of positive and negative charges: K_2SO_4.

Practice Problem 3.6

What are the formulas for all the compounds that can be formed by K^+, Fe^{3+}, Br^-, and SO_4^{2-}?

Further Practice: Questions 3.29 and 3.30 at the end of the chapter

3.4 | NAMING IONIC COMPOUNDS

On their canoe trip, Megan and Jeff took along a packet of instant juice mix. As Jeff mixed the powder with water from his canteen, he noticed the list of ingredients on the label. He asked Megan if she knew what any of the long chemical names meant. Although she didn't recognize them all, several names of compounds were familiar to her. Some were ionic compounds, such as potassium iodide, copper sulfate, and magnesium oxide. Megan knew that they provide essential minerals the body needs for good health. Although she didn't know the formulas for some of the others, she recognized them as sources of other metal ions the body needs. They include ferrous fumarate, which contains iron, and sodium selenite, a source of selenium.

Jeff and Megan, like most new students to chemistry, wondered why chemical names have to be so complicated. Actually, they are not as complicated as they look at first. Chemical names have been established by applying a consistent set of rules that allow chemists to communicate accurately and efficiently. The systematic rules of nomenclature ensure that each chemical compound has a unique name that distinguishes it from all other compounds. Formulas and names for several ionic compounds are listed in Table 3.5. Can you determine the rules used for naming them?

The systematic rules of nomenclature are established by the International Union of Pure and Applied Chemistry (IUPAC).

The names in the table follow from the rules explained in Sections 3.2 and 3.3. Apply them as you name the ionic compounds in Example 3.7.

TABLE 3.5 | Names of Some Ionic Compounds

Formula	Name	Formula	Name
NaCl	sodium chloride	$Mg(NO_3)_2$	magnesium nitrate
$NaNO_2$	sodium nitrite	BaO	barium oxide
$MgCl_2$	magnesium chloride	Li_3N	lithium nitride

EXAMPLE 3.7 | Naming Ionic Compounds

Name the compounds given by the formulas (a) Na_2O and (b) $Ca_3(PO_4)_2$.

Solution:

(a) The first compound, Na_2O, is composed of monatomic cations and anions whose charges can be predicted from the periodic table. We state the name

of the metal ion first, followed by the root of the name for the nonmetal with an *–ide* ending: *sodium oxide.*

(b) The charge on a calcium ion can be predicted from the periodic table. The polyatomic ion is PO_4^{3-}, named phosphate ion. The name of the compound is *calcium phosphate.*

Practice Problem 3.7

Name the compounds K_2O and $MgSO_3$.

Further Practice: Questions 3.39 and 3.40 at the end of the chapter



FIGURE 3.22 Because aluminum has a 3+ charge and oxygen has a 2− charge, the formula for aluminum oxide is Al_2O_3.

If you had trouble establishing rules for naming ionic compounds, the following summary may help. For binary ionic compounds, we give the name of the cation first, simply using its element name. To name the anion we add the suffix *–ide* to the root of the name of the element that corresponds to the anion. Generally for ionic compounds, a prefix indicating the atomic ratio is not used. (As we'll see later in Section 3.5, prefixes are used for molecular compounds.) Thus, NaCl is composed of Na^+ and Cl^- ions and is named sodium chloride. The compound with the formula $AlBr_3$ is composed of Al^{3+} ions and Br^- ions and is named aluminum bromide.

Names for ionic compounds do not directly communicate the relative number of ions in the formula unit. This is unnecessary because we can use the ionic charges for cations and anions predicted from the periodic table or recalled from a table such as Table 3.4 to write the formula unit. For example, the compound aluminum oxide contains an ion of an element that is a metal and an ion of an element that is a nonmetal, so we recognize it as ionic. Aluminum, in group IIIA (13), forms a 3+ ion. Oxygen, in group VIA (16), forms a 2− ion. To be electrically neutral, the formula for aluminum oxide must be Al_2O_3 (Figure 3.22).

Naming ionic compounds containing polyatomic ions is similar to naming binary ionic compounds. The cation is named first, followed by the name of the polyatomic ion. For example, the compound that contains the sodium ion, Na^+, and the sulfate ion, SO_4^{2-}, has the formula Na_2SO_4 and is called sodium sulfate. Some ionic compounds contain the ammonium ion, NH_4^+. An example is ammonium chloride, NH_4Cl. Some rules for naming ionic compounds are summarized in Figure 3.23.

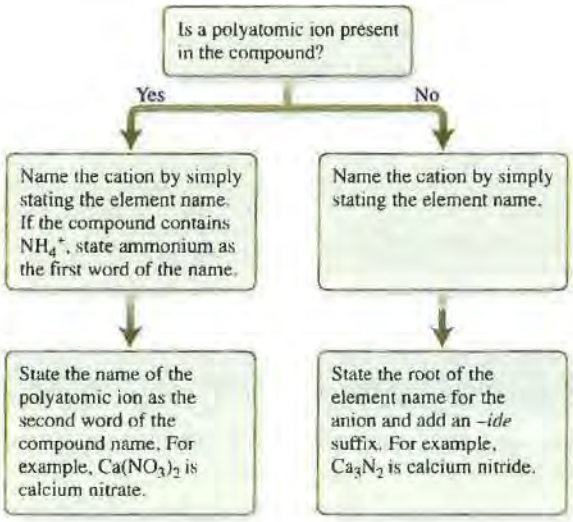


FIGURE 3.23 A flowchart can be used to determine names for some ionic compounds.

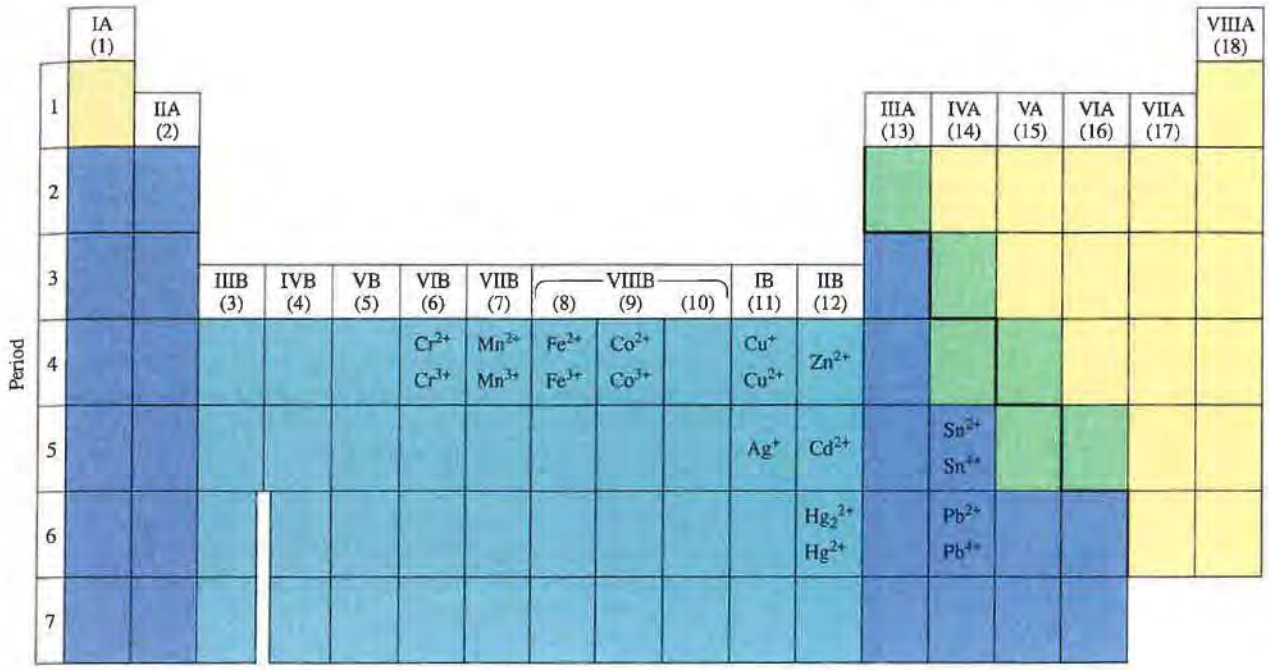


FIGURE 3.24 This periodic table shows some metals that exhibit multiple charges. Notice that zinc, silver, and cadmium are exceptions. They exhibit only one charge.

Some metals can combine with a nonmetal to form more than one ionic compound. For example, copper can combine with chlorine to form CuCl and $CuCl_2$. In these two compounds, the copper has two different charges (1+ and 2+). Many of the metals, especially the transition metals, can exhibit multiple charges (Figure 3.24).

The names we assign to the two compounds of copper and chlorine must be unique so that one can be distinguished from the other. There are two ways of naming compounds of this kind. One is systematic and more current. The other uses common names that were developed—not quite so systematically—in past centuries.

Some formulas and systematic names for ionic compounds containing metals that can have more than one charge are shown in Table 3.6. Try to figure out the rules used to name them.

Did you figure out that the Roman numerals represent charge? Try to use this rule to name the ionic compounds in Example 3.8.

TABLE 3.6 | Ionic Compounds Containing Metals with Multiple Charges

Formula	Cation	Systematic Name
$FeCl_2$	Fe^{2+}	iron(II) chloride
$FeCl_3$	Fe^{3+}	iron(III) chloride
Cu_2O	Cu^{+}	copper(I) oxide
CuO	Cu^{2+}	copper(II) oxide
$CuSO_4$	Cu^{2+}	copper(II) sulfate
SnO	Sn^{2+}	tin(II) oxide

EXAMPLE 3.8 | Naming Ionic Compounds Containing Metals with Multiple Charges

Name the compounds given by the formulas (a) $SnCl_2$ and (b) SnO_2.

Solution:

In both cases the charge on tin is not apparent from its position in the periodic table, so we'll have to deduce its charge in each compound from the anions.

(a) Based on its position in the periodic table, we predict that chloride ion has a 1− charge. For the compound to be electrically neutral, the tin must have a 2+ charge to counter the negative charges from the two chloride ions. We can now write the name for the compound by beginning with the metal name followed by the Roman numeral to represent its charge. Then we write the root of the nonmetal name for the anion with an *–ide* ending. The name of the compound is *tin(II) chloride.*

(b) Based on its position in the periodic table, we predict that oxide ion has a 2− charge. For the compound to be electrically neutral, the tin must have a 4+ charge to counter the four negative charges from the two oxide ions. We can now write the name for the compound by beginning with the metal name followed by the Roman numeral to represent its charge. Then we write the root of the nonmetal name for the anion with an *–ide* ending. The name of the compound is *tin(IV) oxide.*

Practice Problem 3.8

Name the compounds Cr_2S_3 and $FeSO_3$.

Further Practice: Questions 3.43 and 3.44 at the end of the chapter

Zinc, silver, and cadmium normally form only one ion, so Roman numerals are not used when naming their compounds.

If you had trouble establishing rules for naming ionic compounds containing metals that exhibit multiple charges, the following summary may help. The systematic method used to name ionic compounds is called the *Stock system.* It uses Roman numerals in parentheses after the name of the metal to indicate the charge. There is no space between the metal name and the parentheses containing the Roman numeral. For example, CuCl is named copper(I) chloride because the copper in this compound has a charge of 1+. The other compound of copper and chlorine, $CuCl_2$, is named copper(II) chloride because the copper has a 2+ charge. Rules for naming ionic compounds containing a cation that can exhibit multiple charges are shown in Figure 3.25.

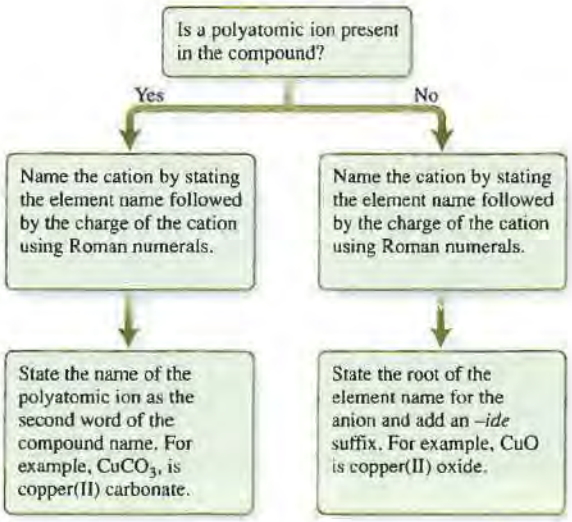


FIGURE 3.25 A flowchart can be used to determine names for ionic compounds that contain metals with multiple charges.

Constructing formulas from Stock system names is particularly straightforward, since the name gives the charge of the metal. For example, in chromium(III) oxide, the chromium has a charge of 3+. From its position in the periodic table, we assume oxygen has a charge of 2−. For the sum of the positive and negative charges to be equal, there must be three oxide ions for every two chromium ions, giving a formula of Cr_2O_3. To practice writing formulas from names, try Example 3.9.

EXAMPLE 3.9 | Formulas for Compounds Containing Metals That Exhibit Multiple Charges

Write the formulas for the compounds with the following names:

(a) iron(III) sulfide
(b) cobalt(II) nitrate

Solution:

(a) The name tells us that iron has a 3+ charge in this compound, Fe^{3+}. Sulfide is the name of the monatomic ion of sulfur that, from its position in the periodic table, we expect to have a charge of 2−, S^{2-}. For the sum of the positive and negative charges to be equal, there must be three sulfide ions for every two iron(III) ions, giving a formula of Fe_2S_3.
(b) We know from the name that cobalt has a 2+ charge in this compound, Co^{2+}. Nitrate is the name of a polyatomic ion with a charge of 1− (NO_3^-). For the sum of the positive charges and the sum of the negative charges to be equal, there must be two nitrate ions for every one cobalt(II) ion, giving a formula of $Co(NO_3)_2$.

Practice Problem 3.9

Write the formulas for the compounds with the following names:

(a) copper(I) sulfate
(b) iron(II) oxide

Further Practice: Questions 3.49 and 3.50 at the end of the chapter

An older method of naming is still commonly used. In this method, the name of the metal is changed by adding a suffix, either *–ous* or *–ic*, to the root of the original Latin name for the metal. (See Chapter 1 for Latin names of some of the elements.) The suffix *–ous* indicates the lower of two charges, and the suffix *–ic* indicates the higher. For example, the two compounds of copper and chlorine described earlier have the formulas CuCl and $CuCl_2$. In the first one, the copper ion has a charge of 1+. This is the lower possible charge, so the compound is called cup*rous* chloride. The other compound, $CuCl_2$, contains copper ions with the higher possible charge, so it is called cupr*ic* chloride. This older method is not often used in laboratories, because it applies only to metals that exhibit two charges. However, it is common on the labels of consumer products. Some examples of common names are listed in Table 3.7, along with the corresponding Stock system names. Example 3.10 asks you to employ the system of common names.

EXAMPLE 3.10 | Common Names for Ionic Compounds Containing Metals That Exhibit More Than One Charge

(a) Using common nomenclature, name the two compounds of iron and oxygen with the formulas FeO and Fe_2O_3.
(b) Write formulas for the compounds cuprous nitrate and cupric nitrate.

Common names are often used on consumer ingredient labels and in medicine. For example, ferrous sulfate, $FeSO_4$, is used as an iron supplement to treat iron-deficiency anemia. This compound has also been used to color concrete. Another compound frequently listed by its common name is stannous fluoride, SnF_2, the source of fluoride in some types of toothpaste.

Solution:

(a) From its position in the periodic table we know that the oxygen ions in both compounds have a charge of 2−. In the first compound, FeO, the iron has a charge of 2+ to make the sum of the positive charges and the sum of the negative charges equal. In the second compound, the iron ion has a charge of 3+. The compound with the formula FeO contains the iron ion with the lower of two possible charges, so it is called *ferrous oxide*. The other compound, Fe_2O_3, contains the iron ion with the higher charge, so it is called *ferric oxide*.

(b) From Figure 3.24 we know that copper ions have two possible charges, 1+ and 2+. The name of the first compound, cuprous nitrate, indicates that it contains copper ions with the lower of two possible charges. This compound contains Cu^+ ions. Nitrate ion has the formula NO_3^-. For the sum of the positive charges and the sum of the negative charges to be equal, the formula for cuprous nitrate must be $CuNO_3$. The name of the second compound, cupric nitrate, indicates that it contains copper ions with the higher of two possible charges, so Cu^{2+} is present in this compound. For the sum of the positive charges and the sum of the negative charges to be equal, the complete formula must be $Cu(NO_3)_2$.

Practice Problem 3.10

(a) Using common nomenclature, name the compounds with the formulas $Sn(SO_4)_2$ and $SnSO_4$.

(b) Write formulas for the compounds ferric sulfide and ferrous sulfide.

Further Practice: Questions 3.51 and 3.52 at the end of the chapter

TABLE 3.7 | Names of Some Ionic Compounds Composed of Metals with Multiple Charges

Formula	Ion	Stock Name	Old Name
$FeCl_2$	Fe^{2+}	iron(II) chloride	ferrous chloride
$FeCl_3$	Fe^{3+}	iron(III) chloride	ferric chloride
Cu_2O	Cu^+	copper(I) oxide	cuprous oxide
CuO	Cu^{2+}	copper(II) oxide	cupric oxide
SnO	Sn^{2+}	tin(II) oxide	stannous oxide
SnO_2	Sn^{4+}	tin(IV) oxide	stannic oxide

3.5 | NAMING AND WRITING FORMULAS FOR MOLECULAR COMPOUNDS

Among the substances Megan and Jeff found in their samples of river water were carbon dioxide and carbon tetrachloride. These are binary molecular compounds. The naming of such compounds follows rules that differ slightly from those used for naming ionic compounds. Several molecular compounds are listed in Table 3.8. Try to determine the rules used to name them.

TABLE 3.8 | Names of Some Common Molecular Compounds

Formula	Name	Formula	Name
CO	carbon monoxide	SO_3	sulfur trioxide
CO_2	carbon dioxide	N_2O_4	dinitrogen tetroxide
CCl_4	carbon tetrachloride	PF_5	phosphorus pentafluoride

Did you figure out how to name binary molecular compounds? Did you determine when and how to use prefixes? Now try to apply the rules to name the molecular compounds in Example 3.11.

EXAMPLE 3.11 | Formulas and Names for Molecular Compounds

Name the compounds (a) P_4O_{10} and (b) NO_2.

Solution:

(a) Since P_4O_{10} is a compound containing two nonmetals, we consider it a molecular compound. Phosphorus is named first, because it occurs farther toward the left and down on the periodic table. Because there are four phosphorus atoms in the molecule, we add the prefix *tetra–*. Oxygen is named second by taking the root of its name and adding the suffix *–ide*. We must add the prefix *deca–* to the beginning of the element name to indicate that there are 10 oxygen atoms in the molecule. The full name is *tetraphosphorus decoxide*. (When two vowels are adjacent, as in deca-oxide, the vowel at the end of the prefix is often dropped.)

(b) The second compound, NO_2, is also molecular, because it contains elements that are nonmetals. Nitrogen is to the left of oxygen, so it is named first. Because there is only one atom of nitrogen in a molecule of the compound, the prefix *mono–* is omitted but implied. The prefix *mono–* is not used for the first element named in a molecular compound. Oxygen is named second by taking the root of its name and adding the suffix *–ide*. We add the prefix *di–* to the beginning of the element name to indicate that there are two oxygen atoms in the molecule. The full name is *nitrogen dioxide*.

Practice Problem 3.11

Name the compounds P_4O_6 and N_2O_5.

Further Practice: Questions 3.61 and 3.62 at the end of the chapter

If you had trouble establishing rules for naming molecular compounds, the following summary may help. Because some nonmetals can combine to form more than one compound, we must identify the appropriate ratio of atoms in the molecule. For example, there are two common compounds containing carbon and oxygen (Figure 3.26). One of them, CO, contains one atom each of carbon and oxygen. Its name is carbon monoxide. (The extra vowel is dropped from mono-oxide.) The other compound, CO_2, contains one atom of carbon and two atoms of oxygen. It is called carbon dioxide.

To name molecular compounds systematically, we begin by stating the name of the element farther to the left and farther down on the periodic table. To the root of the second element's name we add an *–ide* ending. Unlike ionic compounds, where the number of



FIGURE 3.26 Carbon and oxygen form two common molecular compounds.

Molecular compounds are named with prefixes; ionic compounds are not.

TABLE 3.9 | Common Greek Prefixes

Prefix	Number	Prefix	Number	Prefix	Number
mono-	1	penta-	5	octa-	8
di-	2	hexa-	6	nona-	9
tri-	3	hepta-	7	deca-	10
tetra-	4				

The prefix *mono–* is generally not used except in carbon monoxide, CO, and nitrogen monoxide, NO.



FIGURE 3.27 These images represent two molecular compounds containing atoms of sulfur and oxygen.

The formulas SO_3 and SO_3^{2-} correspond to two very different chemical species. The first refers to a compound containing one sulfur atom and three oxygen atoms. The second represents a polyatomic ion that has a 2– charge.

ions in the electrically neutral compound is not stated, we specifically identify the number of atoms in a molecular compound. To indicate this number, we use Greek prefixes such as *mono–*, *di–*, and *tri–*. (These and other prefixes are listed in Table 3.9.)

Let's consider a few compounds to clarify and practice the naming rules. First, consider a compound of sulfur and oxygen in which a molecule contains one sulfur atom and two oxygen atoms, SO_2 (Figure 3.27). Sulfur is listed first in the formula. Similarly, when we name the compound, sulfur will be identified first. Since there is only one sulfur atom in the molecule, a prefix is not necessary. To the root of the name for the second element in the compound, we will add an *–ide* ending, as we did for ionic compounds. We call this compound sulfur dioxide to distinguish it from other compounds containing sulfur and oxygen. Another such compound, SO_3, is shown in Figure 3.27. What is its name?

As another example, consider a compound with the formula HCl. Using the rules just discussed, its name would be hydrogen monochloride. However, the prefix *mono–* is not used because there are no other compounds with this combination of elements. Its name is simply hydrogen chloride. This pattern of naming is similar for HF, HBr, and HI. Rules for naming binary molecular compounds are summarized in Figure 3.28.

Knowing the process by which compounds are named allows us to write formulas for molecular compounds from chemical names. Unlike the formula for an ionic compound, which reveals the simplest ratio of cations and anions in a formula unit, a **molecular formula** tells us the actual number of atoms present in a single molecule of a compound. The formula CO_2, for example, shows that each molecule is made up of one carbon atom and two oxygen atoms. Consider how molecular formulas are written in Example 3.12.

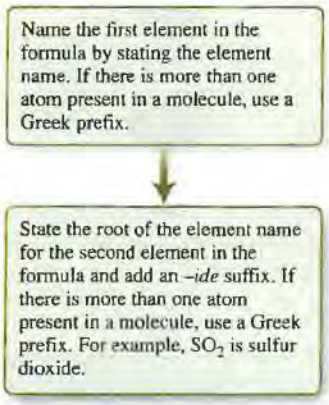


FIGURE 3.28 A flowchart can be used to determine the names for binary molecular compounds.

EXAMPLE 3.12 | Writing Formulas from Names for Molecular Compounds

What is the formula for a compound with the name chlorine dioxide?

Solution:

The first word indicates that chlorine is present in the compound and its symbol should be listed first in the formula. Since there is no prefix preceding its name, one atom is understood. The second element in the formula is oxygen. The prefix *di–* indicates that there are two oxygen atoms in a molecule of this compound. The formula is ClO_2.

Practice Problem 3.12

What is the formula for a compound with the name dinitrogen trioxide?

Further Practice: Questions 3.63 and 3.64 at the end of the chapter

Some binary molecular compounds are known by *trivial,* or nonsystematic, names. Such names often bear no direct relation to the compound's composition. For example, dihydrogen monoxide, H_2O, is commonly known by its trivial name, water. Ammonia, NH_3, and hydrogen peroxide, H_2O_2, are other examples.

In a water molecule, the oxygen atom is the central atom. The formula H_2O is an exception to listing the central atom first in a formula. Other exceptions are H_2S, H_2Se, and H_2Te.

3.6 | ACIDS AND BASES

Megan and Jeff tested their river water samples using an instrument called a pH meter. It revealed that some of their samples tended to be acidic, while others were more basic. An **acid**, simply defined, is a substance that when dissolved in water provides hydrogen ions, H^+. When a compound that is not ionic provides ions when dissolved in water, the process is called *ionization.*

To understand the behavior of acids, consider hydrogen chloride, HCl. By itself this substance is a gas under normal conditions (room temperature and atmospheric pressure). It is a molecular compound that, like other acids, behaves quite differently from the molecular compounds discussed earlier. When gaseous HCl is bubbled into water, it ionizes, producing hydrogen ions, H^+, and chloride ions, Cl^- (Figure 3.29).

The "acid" in acid rain comes from oxides of sulfur and nitrogen, released into the air as gases from the burning of coal and other fuels. The oxides react with water in the air to form sulfuric and nitric acids, which can damage grasslands, forests, and buildings, and threaten the health of humans and other animals.

In general, acids are compounds that usually contain a hydrogen atom that can be removed as an H^+ ion when the compound is placed in water. The process can be summarized for hydrogen chloride by the following equation:

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

A hydrogen ion, H^+, does not actually exist in solution as an isolated entity. Instead, it is surrounded by several water molecules. An H^+ ion in solution is often represented as H_3O^+ (called *hydronium ion*) to show that the H^+ ion is really attached to water molecules.

One important class of acids is the organic (carbon-based) compounds that contain the combination of atoms $-CO_2H$. They are called carboxylic acids. An example is acetic acid, CH_3CO_2H, the substance present in vinegar that gives it its sour taste. The structure of acetic acid is shown in Figure 3.30. Various organic acids are found in the human body and in many of the foods we eat. For example, citric acid is present in citrus fruits such as lemons, oranges, and grapefruit.

We can represent the ionization process in a few ways:

$$HCl(g) \xrightarrow{H_2O} HCl(aq)$$

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

These expressions represent the same process, but the last is the most complete.

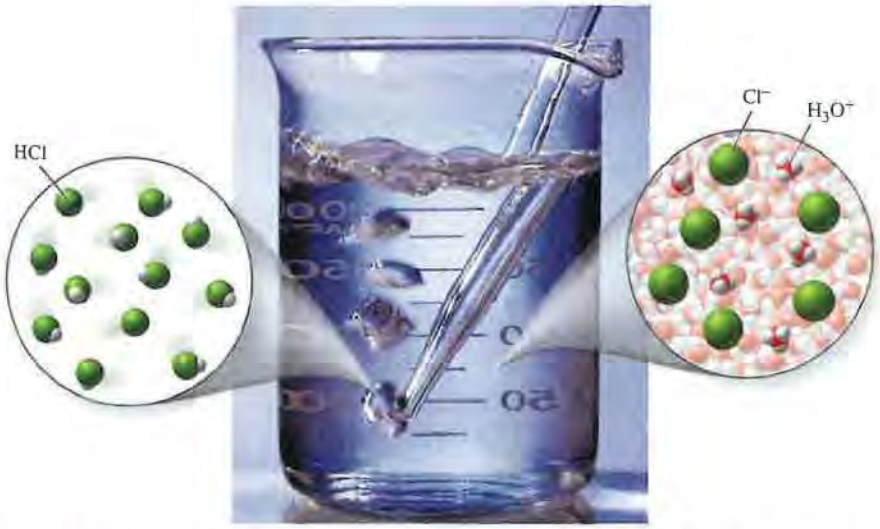


FIGURE 3.29 The molecular compound HCl ionizes to $H^+(aq)$ and $Cl^-(aq)$ when dissolved in water. $H^+(aq)$ is associated with one or more water molecules.

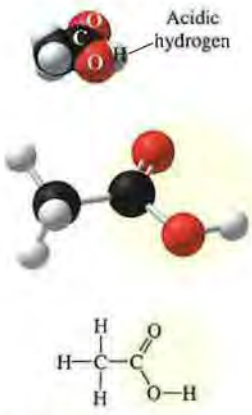


FIGURE 3.30 A class of organic (carbon-containing) compounds contains a $-CO_2H$ group. These compounds are called carboxylic acids. The hydrogen atom attached to an oxygen atom is acidic; the others are not.



FIGURE 3.31 In oxoacids, the acidic hydrogen atoms are attached to oxygen atoms.

Some common bases are:
NaOH, sodium hydroxide
KOH, potassium hydroxide
$Mg(OH)_2$, magnesium hydroxide
NH_3, ammonia

Not all compounds containing hydrogen are acids in water. When hydrogen combines with an element on the far right of the periodic table (excluding the noble gases), the compound is an acid. For example, HF and HCl are both acids, but methane, CH_4, and ammonia, NH_3, are not.

When writing formulas for acids, we generally place the hydrogen first and write (*aq*) after the formula to indicate that the compound is an acid when dissolved in water. For example, HCl(*g*) represents a molecular compound that exists as a gas and we call it hydrogen chloride. The formula HCl(*aq*) corresponds to a solution that's been prepared by dissolving gaseous hydrogen chloride in water and obtaining aqueous hydrogen ions and chloride ions. It's called hydrochloric acid.

Hydrogen combined with a polyatomic ion can also behave as an acid. For example, nitric acid, HNO_3, ionizes in water according to the equation

$$HNO_3(l) \xrightarrow{H_2O} H^+(aq) + NO_3^-(aq)$$

Many other compounds containing hydrogen and oxoanions are acids. Notice in Figure 3.31 that the hydrogen atoms are bound to oxygen atoms in the molecule. When the compound is added to water, the hydrogen atom can separate from the rest of the molecule as an aqueous H^+ ion.

A **base**, simply defined, is a substance that reacts with an acid in aqueous solution to form water. Most common bases either contain hydroxide ion, OH^-, or can provide OH^- ions in solution. For example, sodium hydroxide, NaOH, is an active ingredient in chemical drain cleaners. It is a base that dissolves in water according to the equation

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

Ammonia, NH_3, used in various household cleaners, is another common base, but its behavior is not apparent from its formula. When gaseous ammonia is bubbled into water, a reaction occurs in which some of the ammonia molecules produce aqueous ammonium ions, $NH_4^+(aq)$, and hydroxide ions, $OH^-(aq)$ (Figure 3.32).

Many consumer products contain acids or bases. Some examples are shown in Figure 3.33.

Acids and bases vary in the extent of their ionization or dissociation in water. Strong acids and bases ionize or dissociate completely. They are strong electrolytes.

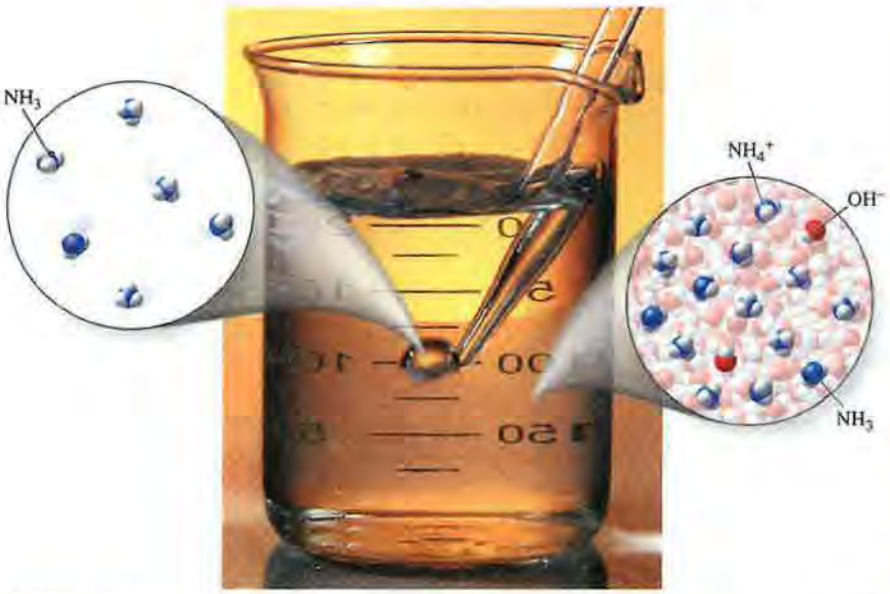


FIGURE 3.32 Some ammonia, NH_3, molecules react with water to produce ammonium ions, NH_4^+, and hydroxide ions, OH^-.

TABLE 3.10 | Some Strong Acids

Formula	Name	Formula	Name
$HCl(aq)$	hydrochloric acid	$H_2SO_4(aq)$	sulfuric acid
$HNO_3(aq)$	nitric acid	$HClO_4(aq)$	perchloric acid

TABLE 3.11 | Names of Some Common Acids

Formula	Name	Formula	Name
$HF(aq)$	hydrofluoric acid	$H_2SO_4(aq)$	sulfuric acid
$HCl(aq)$	hydrochloric acid	$H_2SO_3(aq)$	sulfurous acid
$HI(aq)$	hydroiodic acid	$HClO_4(aq)$	perchloric acid
$H_2S(aq)$	hydrosulfuric acid	$HClO_3(aq)$	chloric acid
$H_2CO_3(aq)$	carbonic acid	$HClO_2(aq)$	chlorous acid
$HNO_3(aq)$	nitric acid	$HClO(aq)$	hypochlorous acid

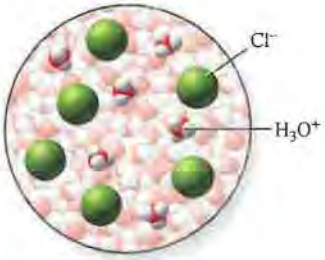


FIGURE 3.33 Many consumer products exhibit acidic or basic properties. For example, window cleaners often contain ammonia and sodas often contain phosphoric acid or citric acid. The next time you are in the grocery store, look at the labels on some products. How can you determine which contain acids or bases?

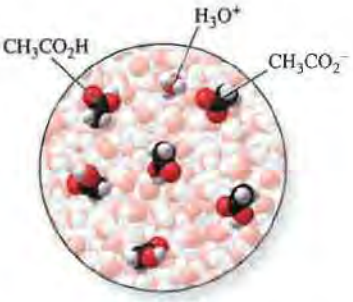
For example, hydrochloric acid is a strong acid used in many industrial and cleaning applications, including some toilet bowl cleaners. It ionizes completely in water to produce ions, so in aqueous solution it is a strong electrolyte. You may encounter several strong acids while working in the laboratory. Some of them are listed in Table 3.10. They must be handled with care, as they can burn skin and eyes and eat holes in clothing. Acetic acid, on the other hand, is present in vinegar and can be consumed safely. In water, relatively few of its molecules ionize. It is both a weak electrolyte and a weak acid. The behaviors of hydrochloric acid and acetic acid are represented in Figure 3.34.

Like the nomenclature of ionic and molecular compounds, the naming of acids follows systematic rules. Table 3.11 shows the formulas and names for many common acids. Can you determine the rules for naming acids?

Did you figure out how to name binary acids? Did you determine how to name acids containing polyatomic ions? Now apply your rules to name the acids in Example 3.13.



Hydrochloric acid



Acetic acid

FIGURE 3.34 Hydrochloric acid ionizes completely. Most acetic acid molecules, on the other hand, remain intact.

EXAMPLE 3.13 | Naming Acids

(a) Name the acid with the formula $HBr(aq)$.
(b) The compound with the formula $HNO_3(aq)$ is called nitric acid. What is the name for the acid with the formula $HNO_2(aq)$?

Solution:

(a) Since $HBr(aq)$ contains just hydrogen and a nonmetal, it is a binary acid. If it were not in aqueous solution, it would be called hydrogen bromide. Since it is in aqueous solution, we can conclude that it is an acid. Binary acids have a *hydro–* prefix, the stem of the element name to which the hydrogen is attached, and an *–ic* suffix. The stem of the name derives from the name of the element hydrogen is attached to. In this case, the element bromine provides the stem *brom.* Assembling the prefix, the stem, and the suffix, along with the word *acid,* gives the name *hydrobromic acid.*

(b) In acids containing a nonmetal that can form two oxoanions, the one with the higher number of oxygen atoms is named with the root for the oxoanion with an *–ic* ending, followed by the word *acid*. Thus, $HNO_3(aq)$ is nitric acid, as stated in the problem. The oxoanion with fewer oxygens is named with the root for the oxoanion name and an *–ous* ending. The name ends with the word *acid*, so $HNO_2(aq)$ is *nitrous acid*.

Practice Problem 3.13

(a) Name the compound $H_2Se(aq)$.

(b) An acid containing an oxoanion of phosphorus with the formula $H_3PO_3(aq)$ is named phosphorous acid. What is the name for the acid with the formula $H_3PO_4(aq)$?

Further Practice: Questions 3.67 and 3.68 at the end of the chapter

Perhaps the rules you came up with allowed you to complete Example 3.13 successfully. If you had trouble, the following summary for naming acids may help. Binary acids in solution are named with the prefix *hydro–* followed by the stem of the name of the nonmetal with the suffix *–ic* and the word *acid* attached. Thus, $HCl(aq)$ is named hydrochloric acid.

Naming acids containing polyatomic ions requires modifying the name of the polyatomic ion. The prefix *hydro–* is not used in naming acids containing polyatomic ions. For example, the compound $H_2CO_3(aq)$, present in many carbonated beverages, contains the carbonate ion. In naming this acid, we remove the *–ate* ending from the name of the polyatomic ion and replace it with *–ic*. Then we add the term *acid* to complete the name. This acid is carbonic acid.

If there is more than one polyatomic ion for a given nonmetal, the names of the acids must distinguish among them. For example, there are two common polyatomic ions containing sulfur. They are the sulfate ion, SO_4^{2-}, and the sulfite ion, SO_3^{2-}. The acids of these ions have the formulas $H_2SO_4(aq)$ and $H_2SO_3(aq)$. In naming the acid containing the sulfate ion, the *–ate* ending is replaced with *–ic* and the term *acid* is added at the end. $H_2SO_4(aq)$ is sulfuric acid. Naming the acid containing the sulfite ion requires us to replace the *–ite* ending with *–ous* and add the term *acid* to the end. Therefore, $H_2SO_3(aq)$ is called sulfurous acid. Rules for naming acids are summarized in Figure 3.35.

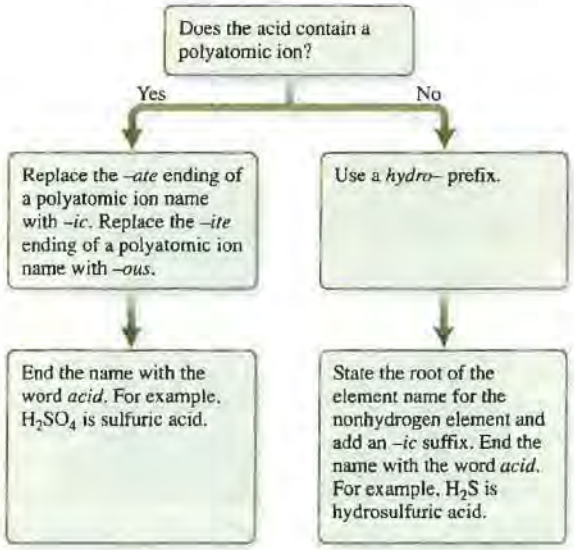


FIGURE 3.35 A flowchart can be used to determine the names for acids.

3.7 | PREDICTING PROPERTIES AND NAMING COMPOUNDS

Jeff and Megan encountered a variety of compounds in their analysis of river water samples. At first they thought that some of the names for compounds were overly complicated, but once they learned the rules, they felt more comfortable. They also realized that by deducing the elements in a compound from its name or formula, they could predict many of its properties. For example, suppose they find out that a compound contains magnesium and chlorine and has the formula $MgCl_2$. Looking at the component elements in the periodic table, they see that the compound contains a metal and a nonmetal. They conclude that the compound is ionic and, therefore, that it probably has some characteristics in common with other ionic compounds. They predict that it might be a brittle solid with a very high melting point. They look up its melting point and find that $MgCl_2$ melts at 712°C. And, because it is an ionic compound, they feel justified in applying the appropriate naming rules and calling it magnesium chloride.

Another ionic compound, $Cr_2(SO_4)_3$, has similar properties. It has the same general properties as $MgCl_2$, and its melting point is even higher, 1857°C. Refer back to Table 3.1, where the general properties of ionic and molecular compounds are listed. Now that we know how to name compounds in these classes, we can predict properties based on their formulas or names. In addition to predicting they have high melting points and boiling points, we would correctly predict the ionic compounds $MgCl_2$ and $Cr_2(SO_4)_3$ would be crystalline solids, be strong electrolytes in water, and conduct electricity when molten or when dissolved in water.

If you don't know which rule to apply to name a particular compound, the flowchart in Figure 3.36 provides some guidance. The questions are designed to help distinguish among ionic compounds, molecular compounds, and acids because they

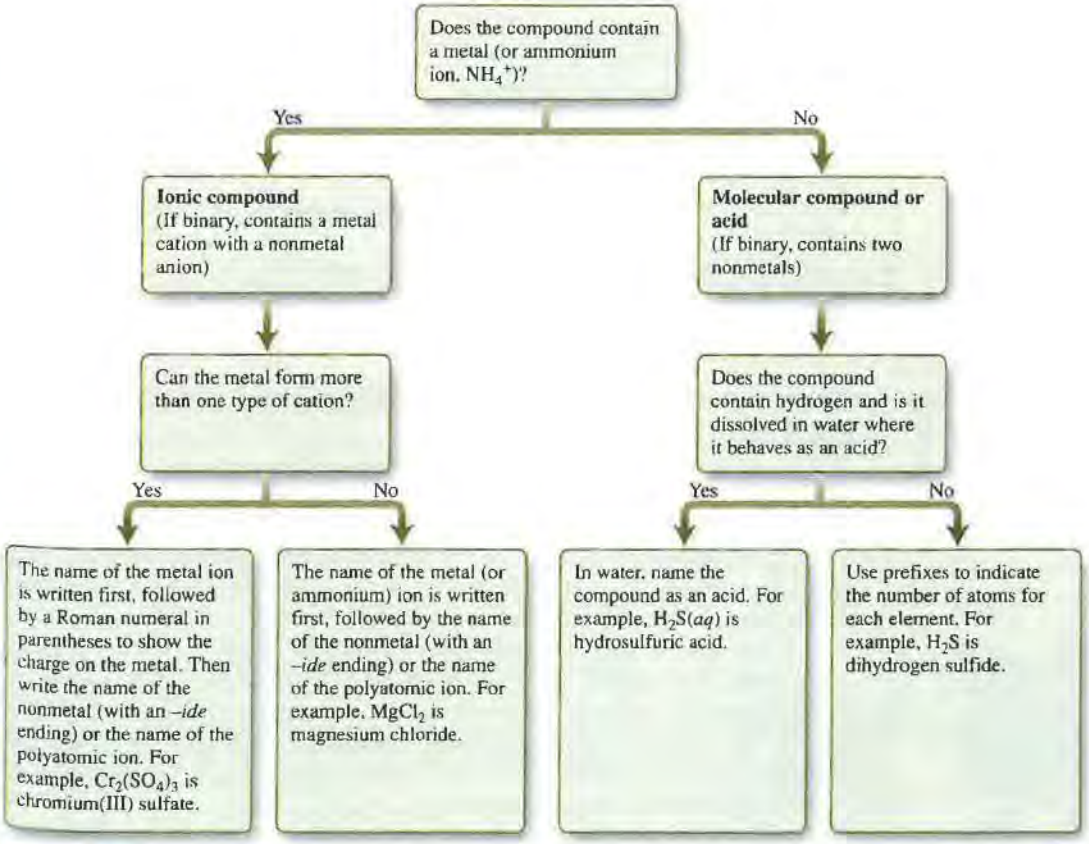


FIGURE 3.36 A flowchart can be used to determine which rules to apply in naming compounds.

Although we use prefixes for molecular compounds (e.g., CCl_4 is carbon tetrachloride), we never use prefixes for ionic compounds (e.g., $AlCl_3$ is aluminum chloride).

are named differently. For example, consider $Cr_2(SO_4)_3$. This compound contains a metal, so we follow the "Yes" branch in response to the first question. It is an ionic compound. The metal exhibits more than one charge, so a Roman numeral is necessary in its systematic name. The metal ion is chromium in the 3+ state, so the first part of the name will be chromium(III). The compound consists of the polyatomic ion SO_4^{2-}, which is the sulfate ion, so the name ends with sulfate. The name of this compound is chromium(III) sulfate.

The compound H_2S does not contain a metal (or ammonium ion), so it is not an ionic compound. It is composed of nonmetals, and its formula does not show that it is dissolved in water, so we will not name it as an acid. From Figure 3.36 we see that it is a molecular compound, named dihydrogen sulfide. From the properties of molecular compounds listed in Table 3.1, we would predict that H_2S has relatively low melting and boiling points. Indeed this is the case, as its melting and boiling points are –85.5°C and –60.7°C, respectively. It is a gas at, and well below, room temperature. When written $H_2S(aq)$, however, this substance is classified as an acid and is named hydrosulfuric acid. We would correctly predict that it is an electrolyte and therefore its solutions conduct electricity. When naming a compound, it is very important to first classify it. Use the flowchart to name the compounds in Example 3.14.

EXAMPLE 3.14 | Using a Flowchart to Name Compounds

Classify and name the compounds with the formulas (a) CCl_4 and (b) $H_3PO_4(aq)$.

Solution:

(a) The compound CCl_4 is molecular because it is composed of two nonmetals. We list the carbon first, followed by the prefix *tetra–* to indicate four chlorine atoms. The name of this compound is *carbon tetrachloride.*

(b) The compound $H_3PO_4(aq)$ is an acid containing the phosphate ion. To name this compound we replace the *–ate* ending with *–ic* and add the word *acid.* This compound is *phosphoric acid.*

Practice Problem 3.14

Classify and name the compounds with the formulas (a) PBr_3, (b) Na_2S, (c) $H_2SO_4(aq)$, and (d) NH_4Cl.

Further Practice: Questions 3.75 and 3.76 at the end of the chapter

SUMMARY

In this chapter we described two broad categories of chemical compounds. Ionic compounds consist of a metal ion (or NH_4^+) combined with a nonmetal ion. Nonmetal ions can be monatomic or polyatomic. Molecular compounds consist of atoms of two or more nonmetals combined in a molecule. Ionic compounds differ from molecular compounds in their physical properties and molecular-level structure.

The ability of a chemical compound to dissociate or ionize in water to provide ions determines how well it conducts electricity. Compounds that dissociate or ionize extensively into ions are strong electrolytes. Examples include soluble ionic compounds, strong acids, and strong bases. Substances, such as weak acids, that ionize only partially in water are weak electrolytes. Those that do not dissociate or ionize when dissolved are nonelectrolytes.

Formulas and names distinguish one compound from all others. The rules for naming compounds were established by international agreement. Formulas for

molecular compounds indicate the number of atoms of each element in a molecule of the compound. Formulas for ionic compounds can only tell us the ratio of cations to anions in a compound.

Within the broad categories of ionic and molecular compounds, some chemical compounds can be described in terms of their acidic or basic properties. Acids are compounds that provide hydrogen ions in solution. Common bases are often compounds that provide hydroxide ions in solution and react with acids to form water.

Compounds are given unique names based on the elements present and how the compound behaves. The rules are summarized in the following table:

Type of Compound	Naming
Ionic	The cation is named first, followed by the name of the anion.
Molecular	The first atom in the formula (the element farther down or to the left in the periodic table) is named first with the second element named as if it were an anion. Greek prefixes are used to designate the number of atoms in the molecule.
Acids	Binary acids are named as *hydro–* followed by the root of the element name with an *–ic* suffix and the word *acid* placed at the end of the name. Acids containing polyatomic ions are named by taking the root of the polyatomic ion name, replacing *–ate* with *–ic* or replacing *–ite* with *–ous* and adding the word *acid* at the end.

KEY TERMS

acid (3.6)
base (3.6)
binary compound (3.1)
electrolyte (3.1)
formula unit (3.3)
ionic compound (3.1)
molecular compound (3.1)
molecular formula (3.5)
monatomic ion (3.2)
nonelectrolyte (3.1)
oxoanion (3.2)
polyatomic ion (3.2)
strong electrolyte (3.1)
weak electrolyte (3.1)

QUESTIONS AND PROBLEMS

The following questions and problems, except for the last section of *Additional Questions*, are paired. Questions in a pair focus on the same concept. Answers to the odd-numbered questions and problems are in Appendix D.

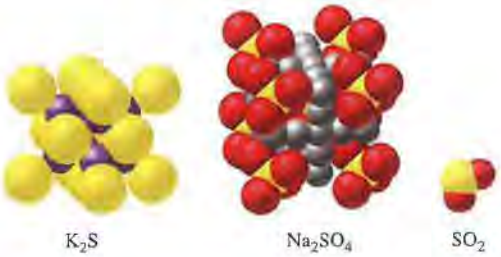
Matching Definitions with Key Terms

3.1 Match the key terms with the descriptions provided.
(a) the smallest repeating unit in an ionic compound
(b) a substance that dissociates or ionizes completely into ions in solution and conducts electricity well
(c) a compound composed of two or more nonmetals that exists in discrete units of atoms held together
(d) a compound that releases hydrogen ions when dissolved in water
(e) a substance that retains its molecular identity when it dissolves and does not conduct electricity
(f) an anion containing oxygen attached to some other element

3.2 Match the key terms with the descriptions provided.
(a) a formula that indicates the actual number of atoms present in a molecule
(b) a substance that dissociates partially into ions upon dissolving and conducts electricity to a slight extent
(c) a compound consisting of cations and anions in proportions that give electrical neutrality
(d) a substance that reacts with an acid to form water
(e) a compound containing atoms or ions of two elements
(f) an ion containing two or more atoms, such as NO_3^-

Ionic and Molecular Compounds

3.3 Identify the following compounds containing sulfur as representations of ionic or molecular compounds.

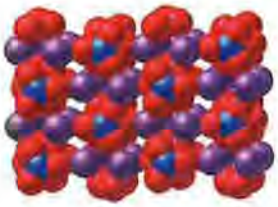


K_2S Na_2SO_4 SO_2

3.4 Identify the following compounds containing nitrogen as representations of ionic or molecular compounds.



Na_3N



KNO_3



NO_2

3.5 Identify the following combinations of elements in a compound as ionic or molecular.
(a) nitrogen and oxygen
(b) potassium and oxygen
(c) phosphorus and fluorine
(d) magnesium and chlorine

3.6 Identify the following combinations of elements in a compound as ionic or molecular.
(a) nitrogen and chlorine
(b) sulfur and oxygen
(c) calcium and chlorine
(d) sodium and nitrogen

3.7 Determine if the following formulas represent ionic or molecular compounds.
(a) PCl_5 (c) $BaCl_2$
(b) LiF (d) N_2O_5

3.8 Determine if the following formulas represent ionic or molecular compounds.
(a) Al_2S_3 (c) $MgBr_2$
(b) P_2O_5 (d) K_2O

3.9 Which of the compounds LiF, CO_2, or N_2O_5 is expected to have the highest melting point?

3.10 Which of the compounds NaCl, CO, or KF is expected to have the lowest melting point?

Monatomic and Polyatomic Ions

3.11 Based on their positions in the periodic table, predict the charge, write the formula, and name the monatomic ions of the following elements.
(a) sodium (b) potassium (c) rubidium

3.12 Based on their positions in the periodic table, predict the charge, write the formula, and name the monatomic ions of the following elements.
(a) fluorine (b) chlorine (c) iodine

3.13 Based on their positions in the periodic table, predict the charge, write the formula, and name the monatomic ions of the following elements.
(a) calcium (b) nitrogen (c) sulfur

3.14 Based on their positions in the periodic table, predict the charge, write the formula, and name the monatomic ions of the following elements.
(a) aluminum (b) barium (c) phosphorus

3.15 The following image represents an oxoanion of nitrogen with a 1− charge. Write a formula for this ion and name it.



3.16 The following image represents an oxoanion of nitrogen with a 1− charge. Write a formula for this ion and name it.



3.17 Name the following polyatomic ions.
(a) SO_4^{2-} (b) OH^- (c) ClO_4^-

3.18 Name the following polyatomic ions.
(a) SO_3^{2-} (b) CN^- (c) ClO_2^-

3.19 Write formulas for the following ions.
(a) nitride (b) nitrate (c) nitrite

3.20 Write formulas for the following ions.
(a) chlorite (b) chloride (c) chlorate

3.21 Write formulas for the polyatomic ions with the following names.
(a) carbonate
(b) ammonium
(c) hydroxide
(d) permanganate

3.22 Write formulas for the polyatomic ions with the following names.
(a) bicarbonate
(b) acetate
(c) periodate
(d) hypochlorite

3.23 This image is a representation for an oxoanion of sulfur with a 2− charge. Write the formula and name this ion.



3.24 This image is a representation for an oxoanion of nitrogen with a 1− charge. Write the formula and name this ion.



3.25 The iodate ion is an oxoanion of iodine bearing a negative one (1−) charge and containing three oxygen atoms. Write the formula for the iodate ion.

3.26 The selenite ion is an oxoanion of selenium bearing a negative two (2−) charge and containing three oxygen atoms. Write the formula for the selenite ion.

Formulas for Ionic Compounds

3.27 Suppose an ionic compound containing aluminum and chloride ions dissolves in water. Beginning with a representation for five aluminum ions, draw a picture of a solution that is electrically neutral.

3.28 Suppose an ionic compound containing magnesium and nitrate ions dissolves in water. Beginning with a representation for five magnesium ions, draw a picture of a solution that is electrically neutral.

3.29 Write the formula of the compound formed by each of the following sets of ions.
(a) Ba^{2+} and Cl^- (c) Ca^{2+} and PO_4^{3-}
(b) Fe^{3+} and Br^- (d) Cr^{3+} and SO_4^{2-}

3.30 Write the formula of the compound formed by each of the following sets of ions.
(a) Na^+ and N^{3-} (c) Ti^{3+} and CO_3^{2-}
(b) Cs^+ and ClO_3^- (d) NH_4^+ and S^{2-}

3.31 Identify the ions in the compounds represented in the following formulas.
(a) KBr (c) $Mg_3(PO_4)_2$
(b) $BaCl_2$ (d) $Co(NO_3)_2$

3.32 Identify the ions in the compounds represented in the following formulas.
(a) NaBr (c) $Ba_3(PO_4)_2$
(b) $AlCl_3$ (d) $Mn(NO_3)_2$

3.33 Iron forms two ions, Fe^{2+} and Fe^{3+}.
(a) What are the formulas for the compounds the two iron ions can form with oxide ion, O^{2-}?
(b) What are the formulas for the compounds the two iron ions can form with chloride ion, Cl^-?

3.34 Two ions of chromium are Cr^{2+} and Cr^{3+}.
(a) What are the formulas for the compounds the two chromium ions can form with oxide ion, O^{2-}?
(b) What are the formulas for the two compounds the chromium ions can form with chloride ion, Cl^-?

3.35 Sodium ion has a 1+ charge: Na^+. In combination with the sulfate ion, it forms the compound Na_2SO_4.
(a) What is the charge on the sulfate ion?
(b) What is the charge on a strontium ion if it forms the compound $SrSO_4$?

3.36 Magnesium has a 2+ charge: Mg^{2+}. In combination with the nitrate ion, it forms the compound $Mg(NO_3)_2$.
(a) What is the charge on the nitrate ion?
(b) What is the charge on a potassium ion if it forms the compound KNO_3?

3.37 The listed formulas are *incorrect*. Determine what is wrong with each and correct it.
(a) $NaCl_2$ (b) KSO_4 (c) Al_3NO_3

3.38 The listed formulas are *incorrect*. Determine what is wrong with each and correct it.
(a) MgCl (b) $Na(SO_4)_2$ (c) $K_2(NO_3)_2$

Naming Ionic Compounds

3.39 Name the following ionic compounds.
(a) $MgCl_2$ (d) KBr
(b) Al_2O_3 (e) $NaNO_3$
(c) Na_2S (f) $NaClO_4$

3.40 Name the following ionic compounds.
(a) $BaCl_2$ (d) KNO_2
(b) $(NH_4)_2S$ (e) $Mg_3(PO_4)_2$
(c) MgO (f) $KClO_2$

3.41 For which of the following compounds must we specify the charge on the metal when we name the compound: $MgSO_4$, $MnSO_4$, $CaCl_2$, $CoCl_2$, $AgNO_3$?

3.42 For which of the following compounds must we specify the charge on the metal when we name the compound: $Cr(NO_3)_2$, $Ba(NO_3)_2$, $AlBr_3$, NiS, $ZnCl_2$?

3.43 Name the following ionic compounds.
(a) Cu_2O (c) $FePO_4$
(b) $CrCl_2$ (d) CuS

3.44 Name the following ionic compounds.
(a) $TiCl_3$ (c) Mn_2O_3
(b) $CoBr_2$ (d) $Fe_3(PO_4)_2$

3.45 Write formulas for the following ionic compounds.
(a) calcium sulfate
(b) barium oxide
(c) ammonium sulfate
(d) barium carbonate
(e) sodium chlorate

3.46 Write formulas for the following ionic compounds.
(a) potassium bicarbonate
(b) sodium oxide
(c) ammonium hydroxide
(d) magnesium bromide
(e) sodium hypochlorite

3.47 Given the following formulas, what is the charge on the metal? What is the name of each compound?
(a) $CoCl_2$ (c) $Cr(NO_3)_3$
(b) PbO_2 (d) $Fe_2(SO_4)_3$

3.48 Given the following formulas, what is the charge on the metal? What is the name of each compound?
(a) $CoBr_3$ (c) Cr_2O_3
(b) $CuCl_2$ (d) $FeSO_4$

3.49 Write formulas for the following ionic compounds.
(a) cobalt(II) chloride
(b) manganese(II) nitrate
(c) chromium(III) oxide
(d) copper(II) phosphate

3.50 Write formulas for the following ionic compounds.
(a) copper(II) bromide
(b) iron(III) nitrate
(c) chromium(VI) oxide
(d) copper(II) chlorate

3.51 What are the common names for $Fe(NO_3)_2$ and $Fe(NO_3)_3$?

3.52 What are the common names for Cu_2SO_4 and $CuSO_4$?

3.53 Complete the following table by writing the names and formulas for the compounds formed when the cations listed across the top combine with the anions listed along the side.

	Ca^{2+}	Fe^{2+}	K^+
Cl^-			
O^{2-}			
NO_3^-			
SO_3^{2-}			
OH^-			
ClO_3^-			

	Mn^{2+}	Al^{3+}	NH_4^+
Cl^-			
O^{2-}			
NO_3^-			
SO_3^{2-}			
OH^-			
ClO_3^-			

3.54 Complete the following table by writing the names and formulas for the compounds formed when the cations listed across the top combine with the anions listed along the side.

	Ba^{2+}	Fe^{3+}	Na^+
Br^-			
S^{2-}			
NO_2^-			
SO_4^{2-}			
HCO_3^-			
ClO_4^-			

	Ni^{2+}	Cr^{3+}	Ag^+
Br^-			
S^{2-}			
NO_2^-			
SO_4^{2-}			
HCO_3^-			
ClO_4^-			

3.55 Complete the following table by writing the formulas for the compounds formed when the cations listed across the top combine with the anions listed along the side.

	potassium	iron(III)	strontium
iodide			
oxide			
sulfate			
nitrite			
acetate			
hypochlorite			

	aluminum	cobalt(II)	lead(IV)
iodide			
oxide			
sulfate			
nitrite			
acetate			
hypochlorite			

3.56 Complete the following table by writing the formulas for the compounds formed when the cations listed across the top combine with the anions listed along the side.

	sodium	chromium(II)	calcium
chloride			
sulfide			
nitrate			
sulfite			
hydroxide			
chlorate			

	ammonium	iron(III)	lead(II)
chloride			
sulfide			
nitrate			
sulfite			
hydroxide			
chlorate			

3.57 What is the formula for silver chloride?

3.58 What is the formula for zinc chloride?

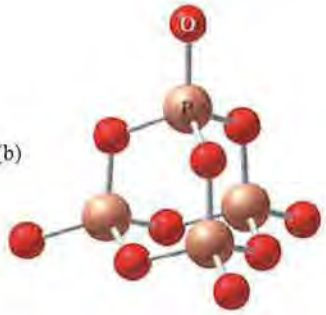
Naming and Writing Formulas for Molecular Compounds

3.59 Write formulas for the molecular compounds represented by the following images.

(a)



(b)



(c)



3.60 Write formulas for the molecular compounds represented by the following images.

(a)



(b)



(c)



3.61 Name the following compounds.
(a) PF_5 (c) CO
(b) PF_3 (d) SO_2

3.62 Name the following compounds.
(a) SO_3 (c) NO_2
(b) N_2O_4 (d) CS_2

3.63 Write formulas for the following compounds.
(a) sulfur tetrafluoride
(b) tricarbon dioxide
(c) chlorine dioxide
(d) sulfur dioxide

3.64 Write formulas for the following compounds.
(a) carbon disulfide
(b) dinitrogen pentoxide
(c) boron nitride
(d) iodine heptafluoride

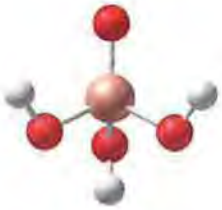
Acids and Bases

3.65 Which of the following images correctly represents phosphoric acid?

A



B



C



3.66 Which of the images shown in Question 3.65 corresponds to phosphorous acid?

3.67 Name the following acids.
(a) HF(*aq*)
(b) $HNO_3(aq)$
(c) $H_3PO_3(aq)$

3.68 Name the following acids.
(a) HI(*aq*)
(b) $HNO_2(aq)$
(c) $H_3PO_4(aq)$

3.69 Write formulas for the following acids.
(a) hydrofluoric acid
(b) sulfurous acid
(c) perchloric acid

3.70 Write formulas for the following acids.
(a) hydroiodic acid
(b) sulfuric acid
(c) hypochlorous acid

3.71 When nitric acid ionizes in water, what ions are present in solution?

3.72 When hydrobromic acid ionizes in water, what ions are present in solution?

Predicting Properties and Naming Compounds

3.73 Which of the following compounds should be crystalline, brittle solids at room temperature and are electrolytes: KI, $Mg(NO_3)_2$, NO_2, NH_3, NH_4NO_3?

3.74 Which of the following compounds should have relatively low melting points and boiling points and be nonconductors of electricity when in the liquid state: CH_3OH, Li_2O, $CaCO_3$, H_2O, NO_2?

3.75 Many sunscreens contain tiny particles (nanoparticles) of either titanium(IV) oxide or zinc oxide. These compounds have the ability to filter a broader range of ultraviolet light than other sunscreening agents. Classify these compounds as ionic or molecular, and write their formulas.

3.76 An injection of ammonium chloride can help correct blood pH levels that are too high (alkaline) due to vomiting or the use of diuretics. Classify this compound as ionic or molecular, and write its formula.

3.77 Carbon dioxide and dinitrogen oxide are two principal greenhouse gases that enter the atmosphere because of human activities. Classify these compounds as ionic or molecular, and write their formulas.

3.78 Sulfur trioxide and nitrogen dioxide gases mix with rainwater to form sulfuric acid and nitric acid. Write the formulas for both gases and both acids.

3.79 Name the compounds represented in Question 3.3.

3.80 Name the compounds represented in Question 3.4.

3.81 Name the following compounds or ions. Identify each compound as molecular or ionic.

(a) NO_3
(b) NO_3^-
(c) KNO_3
(d) Na_3N
(e) $AlCl_3$
(f) PCl_3
(g) TiO
(h) MgO

3.82 Name the following compounds or ions. Identify each compound as molecular or ionic.

(a) SO_4^{2-}
(b) S^{2-}
(c) CO_2
(d) CO_3^{2-}
(e) $CaCO_3$
(f) Fe_2S_3
(g) CS_2
(h) Na_2S

3.83 Write the formula for each of the following.

(a) sodium carbonate
(b) sodium bicarbonate
(c) carbonic acid
(d) hydrofluoric acid
(e) sulfur trioxide
(f) copper(II) sulfate
(g) sulfuric acid
(h) hydrosulfuric acid

3.84 Write the formula for each of the following.

(a) ammonia
(b) ammonium phosphate
(c) hydrochloric acid
(d) potassium hydroxide
(e) chromium(III) oxide
(f) magnesium nitrate
(g) manganese(IV) oxide
(h) nitrogen dioxide

3.85 How do HCl(*aq*) and HCl(*g*) differ? How are they named differently?

3.86 How do $H_2S(aq)$ and $H_2S(g)$ differ? How are they named differently?

3.87 Explain why the name given for each compound is incorrect.

(a) Na_2SO_4, disodium sulfate
(b) CaO, calcium(II) oxide
(c) CuO, copper oxide
(d) PCl_5, phosphorous chloride

3.88 Explain why the name given for each compound is incorrect.

(a) NF_3, nitrogen fluoride
(b) NiS, nickel sulfate
(c) $Ba(NO_3)_2$, barium(II) nitrate
(d) H_2O_2, dihydrogen dioxide

3.89 Explain why the formula given for each compound is incorrect.

(a) potassium sulfide, KS
(b) nickel(II) carbonate, Ni_2CO_3
(c) sodium nitride, $NaNO_3$
(d) nitrogen triiodide, N_3I

3.90 Explain why the formula given for each compound is incorrect.

(a) calcium hydroxide, $CaOH_2$
(b) aluminum chloride, $AlCl_2$
(c) dinitrogen trioxide, N_3O_2
(d) potassium phosphide, K_3PO_4

Additional Questions

3.91 Determine what ions and how many of each are formed for every one formula unit that dissociates in water.

(a) NaCl
(b) $MgCl_2$
(c) Na_2SO_4
(d) $Ca(NO_3)_2$

3.92 Determine what ions and how many of each are formed for every one formula unit that dissociates in water.

(a) KBr
(b) NH_4Cl
(c) K_3PO_4
(d) $Al(NO_3)_3$

3.93 Indicate whether each of the following substances is an electrolyte or a nonelectrolyte.

(a) NaOH(*aq*)
(b) HCl(*aq*)
(c) NaCl(*aq*)
(d) $C_{12}H_{22}O_{11}(aq)$ (sucrose solution)

3.94 Indicate whether each of the following substances is an electrolyte or a nonelectrolyte.

(a) $CH_2Cl_2(aq)$
(b) $HNO_3(aq)$
(c) KBr(*aq*)
(d) KOH(*aq*)

3.95 Many transition metals form more than one ion and are named using Roman numerals to designate the charge on the ion. Why are Roman numerals unnecessary when naming ionic compounds containing silver, zinc, or cadmium?

3.96 Determine what ions are present when each of the following dissociates or ionizes in water.

(a) HCl
(b) NaOH
(c) $CaCl_2$
(d) KNO_3
(e) Na_3PO_4

3.97 Write formulas, including charge, for each of the following polyatomic ions.

(a) nitrate
(b) sulfite
(c) ammonium
(d) carbonate
(e) sulfate
(f) nitrite
(g) perchlorate

3.98 Name the following compounds.
(a) $KBr(s)$
(b) $N_2O_5(g)$
(c) $HBr(aq)$
(d) $Na_2SO_4(s)$
(e) $Fe(NO_3)_3(s)$

3.99 Name the following compounds.
(a) $MgBr_2(s)$
(b) $H_2S(g)$
(c) $H_2S(aq)$
(d) $CoCl_3(s)$
(e) $KOH(aq)$
(f) $AgBr(s)$

3.100 Name the following compounds.
(a) NH_3
(b) H_2O
(c) H_2O_2

3.101 Write formulas for each of the following compounds.
(a) lead(II) chloride
(b) magnesium phosphate
(c) nitrogen triiodide
(d) iron(III) oxide
(e) calcium nitride
(f) barium hydroxide
(g) dichlorine pentoxide
(h) ammonium chloride

3.102 Suppose you need some iron(III) chloride for an experiment. When you go to the stockroom, there is one bottle labeled $FeCl_2$ and another with the formula $FeCl_3$. Which bottle should you select?

3.103 The active ingredient in baking soda is sodium bicarbonate (also called sodium hydrogen carbonate). What is the formula for this compound?

3.104 An active ingredient in toothpaste is tin(II) fluoride. What is the formula of this compound? What is the common name for this compound?

3.105 Calcium hypochlorite is a compound used in water treatment to kill disease-causing organisms. What is the formula of this compound?

3.106 Calcium hydrogen carbonate, magnesium chloride, and calcium sulfate can cause scale buildup in pipes. What are the formulas for these compounds?

3.107 Aluminum sulfate and calcium oxide are used in water treatment to form a compound that allows solid materials to be removed from water.

$$Al_2(SO_4)_3(aq) + 3CaO(s) + 3H_2O(l) \longrightarrow 2Al(OH)_3(s) + 3CaSO_4(s)$$

Based on their formulas, what substances in this chemical equation are molecular?

3.108 Consider the following formulas and names. Explain what is wrong with each name.
(a) $FeBr_2$, iron dibromide
(b) CS_2, copper(IV) sulfide
(c) $Co(NO_3)_3$, cobalt trinitrite
(d) $Mg(OH)_2$, magnesium dihydroxide
(e) Cu_2O, dicopper(II) oxide

3.109 When solid copper metal is added to a solution of silver nitrate, solid silver and copper(II) nitrate form. Write formulas for all the substances involved in this process.

3.110 Indicate whether each of the following combinations of elements would form molecular or ionic compounds.
(a) sodium and iodine
(b) selenium and chlorine
(c) oxygen and chlorine
(d) chromium and oxygen

3.111 Write formulas for the following compounds.
(a) ammonia
(b) nitric acid
(c) nitrous acid

3.112 A sports drink contains the ingredients sucrose, whey protein concentrate, fructose, maltodextrin, citric acid, sodium chloride, lecithin, magnesium oxide, vitamin E, and ferrous sulfate. Write formulas for the molecular and ionic compounds present in this product that can be derived from what you learned in this chapter.

3.113 Methane, CH_4, is not an acid, but CH_3CO_2H is. Given the following figures, which hydrogen is responsible for the acid properties of the second compound, acetic acid?



Methane Acetic acid

3.114 The compounds CH_4 and NH_3 are not acids, but glycine is. Given the structure shown, which hydrogen is responsible for the acidic properties of this amino acid?



3.115 List similarities and differences between each of the following substances.



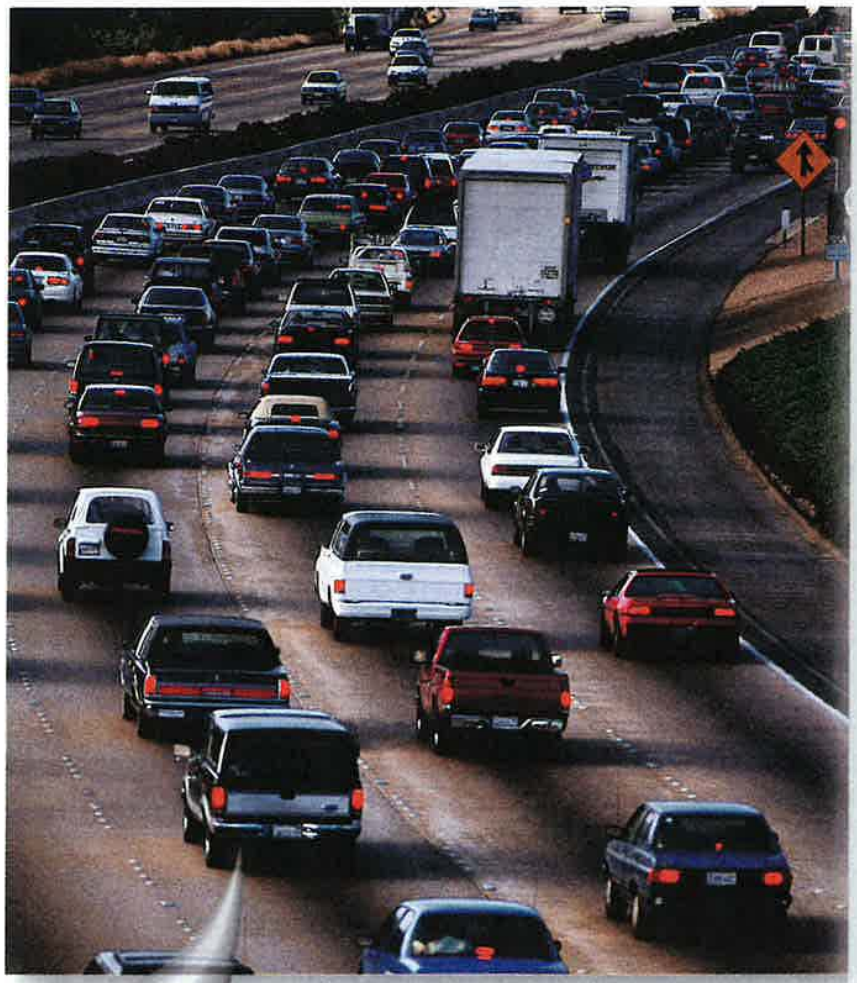
3.116 Write formulas for
(a) potassium peroxide
(b) calcium peroxide

CHAPTER

6

Quantities in Chemical Reactions

6.1 The Meaning of a Balanced Equation
6.2 Mole-Mole Conversions
6.3 Mass-Mass Conversions
6.4 Limiting Reactants
6.5 Percent Yield
6.6 Energy Changes
6.7 Heat Changes in Chemical Reactions
Summary
Key Relationships
Key Terms
Questions and Problems





Imagine the car you will drive five years from now. What will power it? Will it be gasoline? Ethanol? An electric battery? Hydrogen?

Gasoline, a form of petroleum, has been used to power the petroleum-fuel internal-combustion engine (ICE) of automobiles since their invention in the 1860s. Despite steady improvements to the ICE, today's ICE vehicles are only 20% to 25% efficient. This means that 75% to 80% of the energy in gasoline is wasted. This is not the only disadvantage. The combustion of gasoline in the ICE releases pollutants into the atmosphere, and our reliance on gasoline for transportation places a heavy demand on a limited—and rapidly dwindling—natural resource. Although the auto industry has found ways to significantly reduce many toxic gas emissions, the nontoxic carbon dioxide that is released instead is a major factor in the warming of Earth. It is a contributor to the "greenhouse effect," which may be responsible for the recent increase of Earth's average temperature (often called "global warming").

Because Lily is concerned about the environmental impacts of burning fuels, she drives a car that runs on an alternative fuel, an ethanol blend called E-85. It is 85% ethanol (ethyl alcohol, an oxygen-containing fuel) and 15% gasoline. She chose the car because ethanol is a cleaner-burning fuel than gasoline. She's happy that ethanol—unlike gasoline—is a renewable fuel. It can be made from the fermentation of corn or any other carbohydrate. She's proud that she bought her car for about the same price as a gasoline-powered car, but she admits that E-85 isn't always easy to find at service stations.

Joel paid more for his battery-powered electric vehicle (EV), another alternative to the ICE. Instead of filling up with gasoline, Joel recharges his car battery. His engine is about 70% efficient, has fewer mechanical parts than an ICE, and gives off no tailpipe emissions. However, the electricity that recharges Joel's car battery comes from a power plant that burns fossil fuel. The combustion process generates carbon dioxide, so there is still the problem of increased CO_2 in the atmosphere, although it doesn't come directly from Joel's car. Also, the overall efficiency drops to less than 40% when the loss of energy between the power plant and the wall outlet in Joel's garage is taken into account. The drain on scarce fossil fuel resources is no less than using gasoline.

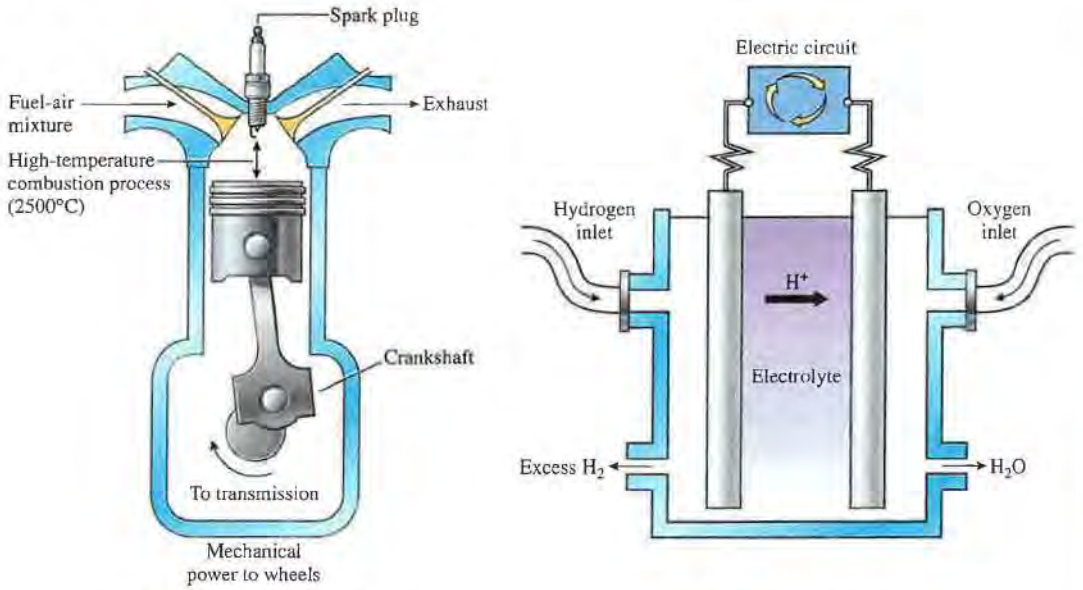
Lily and Joel have heard that hydrogen-powered cars may soon be available. How can hydrogen power a vehicle? When hydrogen reacts with oxygen to form water, a large amount of energy is released. Hydrogen is the major fuel that powers the space shuttle throughout takeoff (Figure 6.1). Hydrogen-powered cars do not burn hydrogen directly; instead, they use a fuel cell, in which the reaction between hydrogen and oxygen is carried out in a controlled way.

A fuel cell works on a different principle than an ICE (Figure 6.2). In an ICE, the fuel is burned and some of the heat energy is converted to work by driving pistons back and forth to run the car. In a hydrogen fuel cell, the hydrogen is not burned. Instead, the cell is separated by a membrane and hydrogen gas runs through one side of the cell, while oxygen runs through the other side. They react together at the membrane, much like a battery. A flow of electrons produces an electric current that runs the vehicle's electric motor. Although the hydrogen fuel cell looks and works like a battery, there is a major difference. The chemicals that run a battery are stored in it, and they eventually run out. In a fuel cell, the hydrogen and oxygen must be continuously delivered to the cell from external storage tanks that can be refilled.

Many people think hydrogen will be the fuel of the future because it is efficient and clean burning. There are, however, problems to overcome. Storing hydrogen gas under high pressure in a vehicle is a safety hazard. Also, hydrogen gas is explosive in air when not controlled. Solutions to these problems have focused on storing the hydrogen as part of a metal hydride compound such as TiH_3, or in methanol, CH_3OH. Hydrogen can then be obtained from chemical reactions of these compounds



FIGURE 6.1 The reaction between hydrogen and oxygen produces the energy to propel the space shuttle into space. The product of this combustion reaction, H_2O, is visible as clouds of steam.



A Single-cylinder internal-combustion engine (ICE)

B Hydrogen fuel cell

FIGURE 6.2 (A) In an internal-combustion engine (ICE), a fuel-air mixture burns inside a combustion chamber, providing the force needed to turn a crankshaft. (B) A hydrogen fuel cell works something like a battery. A controlled reaction creates a current that runs the vehicle's electric motor.

as needed. Another problem is that hydrogen is not readily available. It is most easily obtained from reactions of methanol or other fuels, which also form carbon monoxide or carbon dioxide. In this case, no reduction of air pollution occurs.

In Iceland, where electricity is produced from geothermal power, electrolysis is used to produce hydrogen, and hydrogen fuel cells are used to run buses. As a result of the success in Iceland, other countries are piloting hydrogen-fueled vehicles (Figure 6.3).



FIGURE 6.3 This prototype hydrogen-fueled car has its fuel cell visible in the left side of the opened engine compartment.

As drivers of alternative-fuel vehicles, Lily and Joel are interested in the amount of fuel they need for their vehicles and the energy output provided by the fuel. Whatever the chemical reaction may be, it is often important to know the quantities of the reactants and products and the amount of energy involved in a reaction. For example, we might want to know the amount of oxygen needed to react with the hydrogen that powers a space-shuttle liftoff, or we might want to produce a certain amount of hydrogen from a reaction of methanol. In another situation, we might be interested in producing a certain amount of energy in a combustion reaction.

In this chapter we will look at the quantitative aspects of chemical reactions. We will show how to determine the amounts of reactants that are needed and the amounts of products that should be formed in a given reaction. We will investigate the amounts of energy produced in chemical reactions, such as the combustion of fuels, and we will compare the amount of heat released by different fuels. While combustion reactions release energy, some chemical reactions absorb energy. We will look at how to measure the amount of energy released or absorbed during chemical reactions.

Questions for Consideration

6.1 What do the coefficients in balanced equations represent?

6.2 How can we use a balanced equation to relate the number of moles of reactants and products in a chemical reaction?

6.3 How can we use a balanced equation to relate the mass of reactants and products in a chemical reaction?
6.4 How do we determine which reactant limits the amount of product that can form?
6.5 How can we compare the amount of product we actually obtain to the amount we expect to obtain?
6.6 How can we describe and measure energy changes?
6.7 How are heat changes involved in chemical reactions?

Math Tools Used in This Chapter

Significant Figures (Math Toolbox 1.2)
Units and Conversions (Math Toolbox 1.3)

6.1 | THE MEANING OF A BALANCED EQUATION

As you learned in Chapter 5, a balanced equation represents the identity and physical state of the reactants and products in a chemical reaction. In addition the coefficients in the equation show the relative amounts of reactants that combine and the relative amounts of products expected to form. Let's consider the combustion reaction of propane, a chemical reaction that occurs when you light a gas grill. The balanced equation for the combustion of propane (Figure 6.4) shows that when 1 propane, C_3H_8, molecule reacts, 5 O_2 molecules combine with it to form 3 CO_2 molecules and 4 H_2O molecules:

Propane for barbecue grills is stored under pressure in its liquid state. Air, which is about 20% oxygen, provides the oxygen for the reaction.

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

Since the coefficients in this equation represent the *relative* numbers of reactant and product molecules, the equation reveals the amounts of reactants and products at any scale. For example, if 2 molecules of C_3H_8 react, they must combine with 10 molecules of O_2, producing 6 molecules of CO_2 and 8 molecules of H_2O. The relative numbers of reactants and products remain the same (1 C_3H_8 : 5 O_2 : 3 CO_2 : 4 H_2O).

Suppose 100 molecules of C_3H_8 react, or suppose 6.022×10^{23} molecules (a mole of molecules) react. How many O_2 molecules will combine with each of these quantities? How many molecules of CO_2 and H_2O should form in each case? Compare your answers to those shown in Table 6.1.

Because 6.022×10^{23} molecules equals 1.000 mol of molecules, we can also say that 1.000 mol of C_3H_8 reacts with 5.000 mol of O_2 to produce 3.000 mol of CO_2

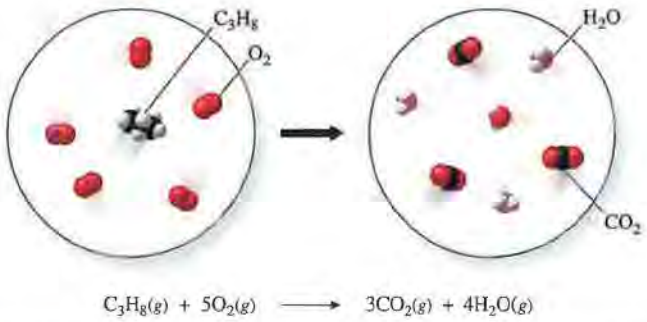


FIGURE 6.4 The balanced equation for the combustion of propane shows the relative numbers of propane and oxygen molecules that react and the relative numbers of carbon dioxide and water molecules that are produced.

TABLE 6.1 | Relationships between Reactants and Products

$C_3H_8(g)$	+	$5O_2(g)$	⟶	$3CO_2(g)$	+	$4H_2O(g)$
1 molecule		5 molecules		3 molecules		4 molecules
2 molecules		10 molecules		6 molecules		8 molecules
100 molecules		500 molecules		300 molecules		400 molecules
6.022×10^{23} molecules		$5 \times (6.022 \times 10^{23})$ molecules		$3 \times (6.022 \times 10^{23})$ molecules		$4 \times (6.022 \times 10^{23})$ molecules
1.000 mol		5.000 mol		3.000 mol		4.000 mol

and 4.000 mol of H_2O. A balanced equation, therefore, also describes the *relative number of moles* of reactants and products. Table 6.1 summarizes the relationships between reactants and products for the combustion of propane.

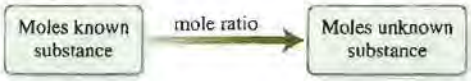
If we know the amount of any reactant or product, we can determine the amount of any other reactant or product in the same reaction by using the balanced equation. The process of determining the amounts of substances in a chemical reaction is called **stoichiometry**. In the next sections, we will use stoichiometry to relate amounts of reactants and products using balanced equations.

6.2 | MOLE-MOLE CONVERSIONS

A convenient way to relate the moles of a reactant or a product to other reactants or products is to use *mole ratios*. Mole ratios are obtained from the coefficients in a balanced equation. Let's look again at the balanced equation for the combustion of propane:

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

The mole ratio of O_2 to C_3H_8 is 5/1, or $\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}$, because the coefficient in front of O_2 is 5 and the coefficient in front of C_3H_8 is an implied 1. This mole ratio tells us that 5 mol of O_2 are needed for every 1 mol of C_3H_8 that reacts. The same equation also shows that the mole ratio of the product CO_2 to the reactant O_2 is 3/5, or $\frac{3 \text{ mol } CO_2}{5 \text{ mol } O_2}$. Mole ratios help us determine the moles of one substance in a reaction when the number of moles of another substance in a reaction is known:

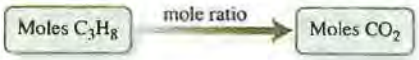


Suppose we want to know the amount of CO_2 that will be produced when 9.21 mol of C_3H_8 reacts? What mole ratio do we use? From the coefficients in the equation, the mole ratio relating CO_2 and C_3H_8 can be written in two ways:

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} \quad \text{and} \quad \frac{1 \text{ mol } C_3H_8}{3 \text{ mol } CO_2}$$



We can use these mole ratios as conversion factors. (See Math Toolbox 1.3 for a review of unit conversions.) We multiply the known moles of C_3H_8 by the ratio that will cancel the old units (mol C_3H_8) and introduce the new units (mol CO_2):



$$\text{mol } CO_2 = 9.21 \cancel{\text{ mol } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \cancel{\text{ mol } C_3H_8}} = 27.6 \text{ mol } CO_2$$

Example 6.1 gives you the opportunity to practice using mole ratios as conversion factors.

EXAMPLE 6.1 | Mole-Mole Conversions

If 1.14 mol of CO_2 was formed by the combustion of C_3H_8, how many moles of H_2O were also formed?

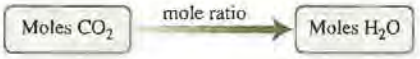
$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

Solution:

We know the number of moles of CO_2, and we want to know the moles of H_2O:



The relationship we use to convert from moles of CO_2 to moles of H_2O is the mole ratio we get from the balanced equation:



First we must ensure that the equation is balanced. Yes it is, so the coefficients in the equation give mole relationships between CO_2 and H_2O, which can be written in two ways:

$$\frac{3 \text{ mol } CO_2}{4 \text{ mol } H_2O} \quad \text{and} \quad \frac{4 \text{ mol } H_2O}{3 \text{ mol } CO_2}$$

The mole ratio to be used as the conversion factor is selected so that multiplying the known moles of CO_2 by the ratio will cancel the old units (mol CO_2) and introduce the new units (mol H_2O):

$$\text{mol } H_2O = 1.14 \cancel{\text{ mol } CO_2} \times \frac{4 \text{ mol } H_2O}{3 \cancel{\text{ mol } CO_2}} = 1.52 \text{ mol } H_2O$$



Notice that the units cancel properly. We would expect the moles of H_2O to be greater than the moles of CO_2, based on the 4:3 ratio in the balanced equation, so this answer makes physical sense.

Practice Problem 6.1

Pure methanol is used as a fuel for all race cars in the Indy Racing League and in the Championship Auto Racing Teams. It is used because methanol fires are easier to put out with water than the fires of most other fuels. The balanced equation for the combustion of methanol is

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(g)$$

Given that 1.00 gal of methanol contains 94.5 mol, how many moles of oxygen gas will react with 1.00 gal of methanol?

Further Practice: Questions 6.11 and 6.12 at the end of the chapter

6.3 | MASS-MASS CONVERSIONS

Among the principles of green chemistry are the prevention of waste and the maximization of atom economy. These goals can be accomplished by careful calculation of the exact amounts of reactants needed to produce the desired amount of product in a chemical reaction.

When we measure the amount of a reactant or product, we do not measure moles because our measuring devices do not count moles. Instead, we typically measure the mass of a reactant or product, but a chemical equation does not directly tell us the *mass* relationship between reactants and products. It tells us the *mole* relationship. This is not a problem, however, because we can convert from mass units (grams) to moles using the molar mass (*MM*) of the substance, as you learned in Chapter 4. Once we convert to moles, we can determine the moles of another substance in the equation using the mole ratio as we did in Section 6.2. Then, if we want to know the mass of the substance, we can convert to grams using its molar mass. The process always involves the same general conversion steps illustrated in Figure 6.5.

To illustrate a mass-mass conversion, let's consider the combination reaction of sodium metal and chlorine gas to form sodium chloride:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Although chemicals are often measured by mass, some substances are more conveniently measured by volume. If the substance is pure and the density is known, then the mass can be determined from the density and volume using the relationship you learned in Chapter 1:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{Mass} = \text{density} \times \text{volume}$$

Sometimes reactants are in the gaseous state or are in solution. Calculations involving these types of substances will be discussed in Chapters 9 and 11.

Suppose we have 9.20 g of sodium metal. How many grams of chlorine gas should react with this amount of sodium in this reaction? Let's follow the general procedure for a mass-mass problem shown in Figure 6.5. Because the balanced equation does not give us a relationship between the grams of Na and the grams of Cl_2, we must first convert grams of Na to moles of Na using its molar mass, 22.99 g/mol.



The molar mass of Na gives the conversion factor between grams of Na and moles of Na, which can be written in two ways:

$$\frac{1 \text{ mol Na}}{22.99 \text{ g Na}} \quad \text{and} \quad \frac{22.99 \text{ g Na}}{1 \text{ mol Na}}$$

When we convert from grams to moles, we multiply the known amount in grams by the ratio that will cancel the old units (grams) and introduce the new units (moles):

$$\text{Moles of Na} = 9.20 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.400 \text{ mol Na}$$

MATH TOOLBOX 1.3

Now that we know the moles of the reactant, Na, we can determine the moles of Cl_2 using the mole ratio of Cl_2 to Na obtained from the balanced equation as the conversion factor:



We multiply the known moles of Na by the ratio that will cancel the old units (mol Na) and introduce the new units (mol Cl_2):

$$\text{Moles of } Cl_2 = 0.400 \text{ mol Na} \times \frac{1 \text{ mol } Cl_2}{2 \text{ mol Na}} = 0.200 \text{ mol } Cl_2$$

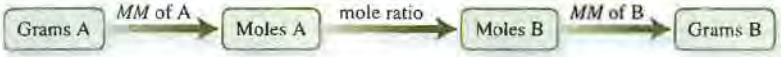


FIGURE 6.5 The process for converting from the mass of one reactant or product to another involves three basic steps.

Once we know the moles of Cl_2, all we have left to do is convert from moles to grams using the molar mass of Cl_2, 70.90 g/mol.



We multiply 0.200 mol Cl_2 by the mole ratio that cancels the old units (moles) and introduces the new units (grams):

$$\text{Mass of } Cl_2 = 0.200 \ \cancel{\text{mol } Cl_2} \times \frac{70.90 \text{ g } Cl_2}{1 \ \cancel{\text{mol } Cl_2}} = 14.2 \text{ g } Cl_2$$

MATH TOOLBOX 1.2

Notice that the answer has three significant figures because the mole amount of Cl_2 has three. (See Math Toolbox 1.2.) When we convert from moles to grams, we always *multiply* by the molar mass.

Alternatively we can do all three conversions consecutively in one calculation:



$$\text{Mass of } Cl_2 = 9.20 \ \cancel{\text{g Na}} \times \frac{1 \ \cancel{\text{mol Na}}}{22.99 \ \cancel{\text{g Na}}} \times \frac{1 \ \cancel{\text{mol } Cl_2}}{2 \ \cancel{\text{mol Na}}} \times \frac{70.90 \text{ g } Cl_2}{1 \ \cancel{\text{mol } Cl_2}} = 14.2 \text{ g } Cl_2$$

All mass-mass conversion problems involve the same three-step calculation. Writing the conversion calculation with all steps combined has advantages. It is easier to ensure that the units cancel properly. Also, you are less likely to make mistakes.

This calculation shows that 9.20 g of Na should react with 14.2 g of Cl_2. Now let's determine how much NaCl should form. Since we know the masses of all but one substance in the reaction, NaCl, we can use the *law of conservation of mass* to determine the mass of NaCl. The mass of the reactants must equal the mass of the products (Figure 6.6):

$$\text{Mass of reactants} = \text{mass of products}$$
$$\text{Mass Na} + \text{mass } Cl_2 = \text{mass NaCl}$$
$$9.20 \text{ g Na} + 14.2 \text{ g } Cl_2 = 23.4 \text{ g NaCl}$$

In this example, we could calculate the mass of NaCl in this way because we knew the mass of all the other substances in the reaction. Usually this is not the case, and we have to use the mass-mass conversion procedure to determine the mass of a reactant or product from a known mass of another reactant or product. Try calculating the mass of NaCl from the mass of either Na (9.20 g) or Cl_2 (14.2 g) using mass-mass conversions. You should get the same answer (23.4 g NaCl) as you would by summing the masses of the reactants. Example 6.2 gives you practice converting from the mass of one reactant or product to the mass of another.

The way you round numbers will affect your answer. You'll get one result if you round after each step, but another if you hold the answers to each step in your calculator. Don't be alarmed if your answers are *slightly* different from those given in this textbook. We generally carry excess digits through the entire calculation, but round the answer to each step to the appropriate number of significant figures.



9.20 g Na (0.400 mol) $2Na(s)$ + 14.2 g Cl_2 (0.200 mol) $Cl_2(g)$ ⟶ 23.4 g NaCl (0.400 mol) $2NaCl(s)$

FIGURE 6.6 The law of conservation of mass states that the mass of the reactants must equal the mass of the products. Mass is not gained or lost. (The amounts given do not represent the quantities of substances shown.)

EXAMPLE 6.2 | Mass-Mass Conversions

The decomposition of nitroglycerin is an explosive reaction that produces four different gases. What mass of water vapor should be produced when 1.0 g of nitroglycerin decomposes by the following reaction?

$$4C_3H_5O_9N_3(l) \longrightarrow 12CO_2(g) + 6N_2(g) + O_2(g) + 10H_2O(g)$$

Solution:

Because we are given grams of nitroglycerin, $C_3H_5O_9N_3$, we must first convert to moles using its molar mass, 227.10 g/mol:

Grams $C_3H_5O_9N_3$ —(*MM* of $C_3H_5O_9N_3$)→ Moles $C_3H_5O_9N_3$

$$\text{Moles of } C_3H_5O_9N_3 = 1.0 \text{ g } C_3H_5O_9N_3 \times \frac{1 \text{ mol } C_3H_5O_9N_3}{227.10 \text{ g } C_3H_5O_9N_3}$$

$$= 0.0044 \text{ mol } C_3H_5O_9N_3$$

MATH TOOLBOX 1.3

Next we use the mole ratio of H_2O to $C_3H_5O_9N_3$ to convert to moles of H_2O:

Moles $C_3H_5O_9N_3$ —(mole ratio)→ Moles H_2O

$$\text{Moles of } H_2O = 0.0044 \text{ mol } C_3H_5O_9N_3 \times \frac{10 \text{ mol } H_2O}{4 \text{ mol } C_3H_5O_9N_3}$$

$$= 0.011 \text{ mol } H_2O$$

Now that we know the moles of H_2O, all we have left to do is convert moles to grams by multiplying by the molar mass of H_2O. Note that the units cancel properly.

Moles H_2O —(*MM* of H_2O)→ Grams H_2O

$$\text{Mass of } H_2O = 0.011 \text{ mol } H_2O \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.20 \text{ g } H_2O$$

Alternatively, we can do all three conversions consecutively in one calculation:

$$\text{Mass of } H_2O = 1.0 \text{ g } C_3H_5O_9N_3 \times \frac{1 \text{ mol } C_3H_5O_9N_3}{227.10 \text{ g } C_3H_5O_9N_3} \times$$

$$\frac{10 \text{ mol } H_2O}{4 \text{ mol } C_3H_5O_9N_3} \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.20 \text{ g } H_2O$$

When 1.0 g of nitroglycerin decomposes, 0.20 g of water vapor is produced.

Practice Problem 6.2

Hydrogen and oxygen are used as fuel for liftoff during a space-shuttle launch. The hydrogen and oxygen are stored within the shuttle as compressed liquids. During an average liftoff, 180,000 kg of hydrogen burns as it reacts with oxygen to form water vapor. The balanced equation for this reaction is

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

(a) What mass of oxygen gas is consumed?
(b) What mass of water vapor is produced?
(c) Does the mass of water vapor equal the sum of the masses of the reactants? What law is followed here?

Further Practice: Questions 6.21 and 6.22 at the end of the chapter

ANIMATION: Limiting Reagent

6.4 | LIMITING REACTANTS

When Lily burns ethanol in her car, she relies on oxygen in the air to support the combustion. She doesn't worry about running out of oxygen. A chemist would say that oxygen is in excess for the reaction.

When the space shuttle is in space, however, there is no air to provide oxygen. The shuttle must take the right amounts of hydrogen and oxygen with it. If too much of one reactant were transported, the extra reactant would have nothing to react with. It would simply add extra mass to the shuttle. Mass-mass calculations are important in such situations.

When a space shuttle takes off, it leaves with measured amounts of hydrogen and oxygen so that no extra mass is added. Many reactions, however, occur with an excess of one reactant. Consider a single-displacement reaction between magnesium metal and aqueous hydrochloric acid:

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

How do we know if all of each reactant is used up in the reaction? If one of the reactants does not react completely, there will be some of it left over at the end of the reaction. The leftover amount will be present mixed with the product, as shown in Figure 6.7. In which of the pictures of Figure 6.7 is magnesium left over? In which is hydrochloric acid left over?

When there is excess magnesium, we see the magnesium metal remaining with the solution. When there is excess hydrochloric acid, no metal remains with the product solution, which also contains excess acid. If we do not want any unreacted magnesium or hydrochloric acid left over, we should mix the reactants in the proper quantities so that their mole ratio is the same as that shown by the coefficients in the balanced equation:

$$\frac{1 \text{ mol Mg}}{2 \text{ mol HCl}}$$

Why don't we use the exact amounts of each reactant that are needed for the reaction? Small errors can always lead to a slight excess of one reactant that might contaminate the product. To ensure that a selected reactant is used completely, we might use an excess of the other reactant. For example, silver bromide used in photographic films is prepared by the reaction of magnesium bromide and silver nitrate solutions. Magnesium bromide is used in excess to ensure that all of the more expensive silver nitrate is used completely.

$HCl(aq)$ can be detected in the product solution with various acid test strips.



A B C

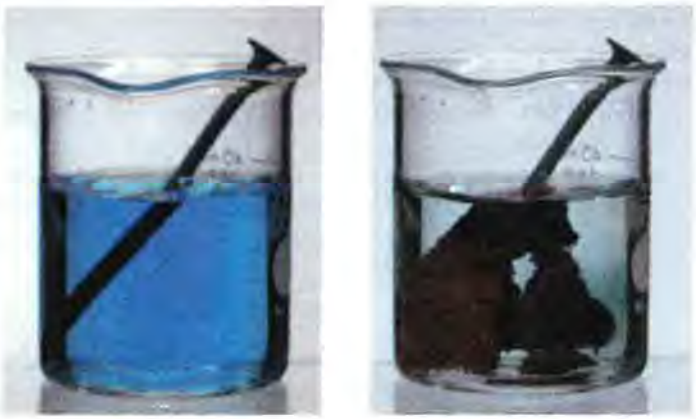
FIGURE 6.7 When reactants are not mixed in relative amounts as described by the balanced equation, one reactant does not react completely. (A) Here magnesium metal reacts with hydrochloric acid to form magnesium chloride in solution and hydrogen gas. When the reaction is complete, one reactant remains in excess. What is the excess reactant in beaker B? In beaker C?

If the reactants mixed together are *not* present in the exact mole ratio, then we will have too much of one reactant and not enough of another. One reactant reacts completely, and some of the other is left unreacted. The reactant that reacts completely is called the **limiting reactant**; it limits the amount of the other reactant that can react, and it limits the amount of product that can form. A reactant that does not react completely is *in excess*.

A limiting reactant is also known as a limiting reagent.

EXAMPLE 6.3 | Identifying the Limiting Reactant

Iron metal is added to an aqueous solution of copper(II) chloride. The solution is blue because copper(II) ions in solution produce a blue color. When the reaction is complete, the solution is colorless. The piece of iron is coated with a granular, brownish-black metal. Identify the limiting reactant and the reactant that is in excess.



Solution:

The loss of all color from the solution indicates that the copper(II) chloride reacted completely. It is the limiting reactant. The iron metal is in excess because some remains unreacted.

Practice Problem 6.3

In Example 6.3 you identified the limiting reactant. Describe what you would observe if the other reactant was the limiting reactant.

Further Practice: Questions 6.29 and 6.30 at the end of the chapter

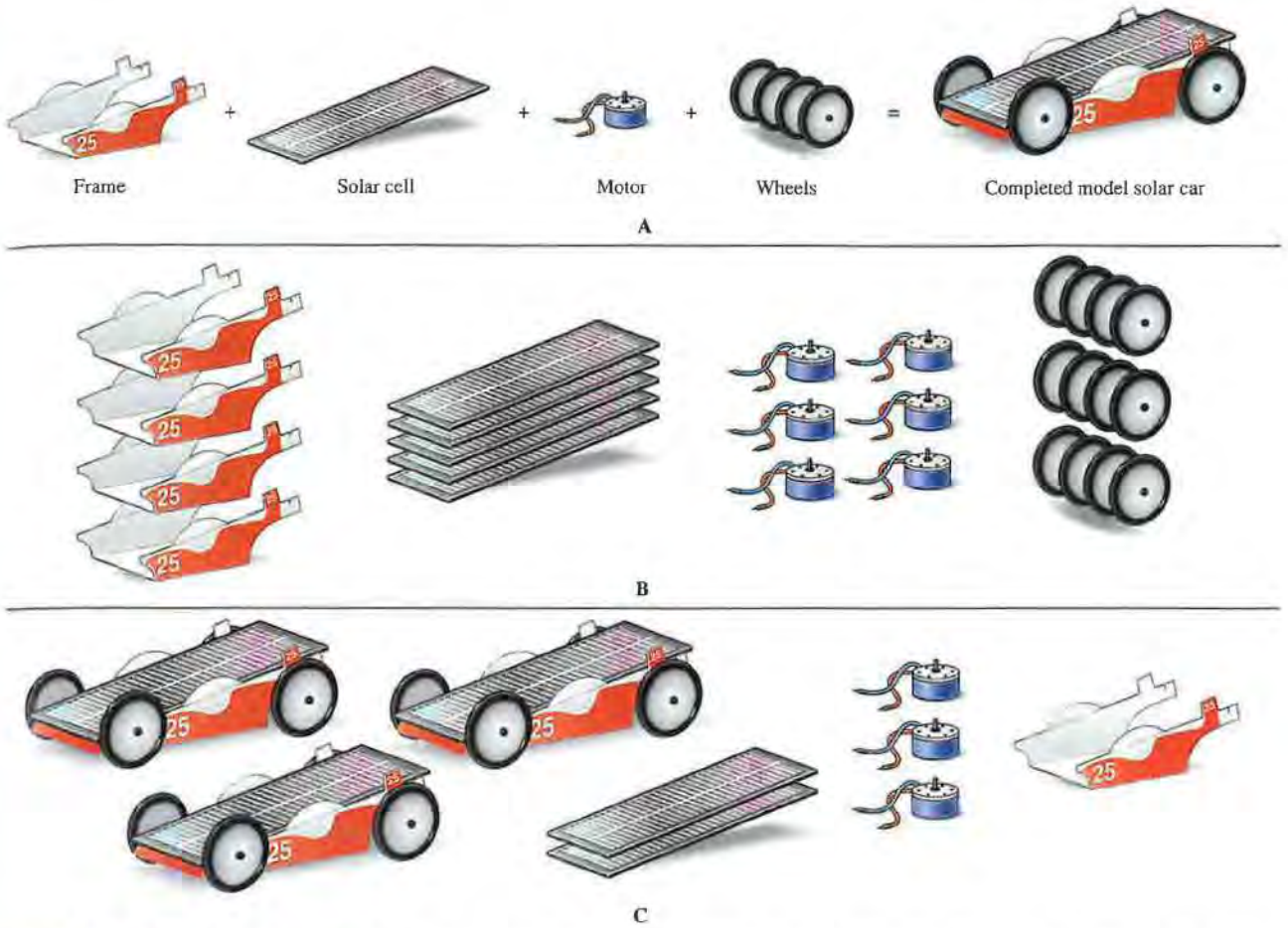


FIGURE 6.8 To make a model solar car, a specific number of each part is required. To make more than one car, the relative number of parts remains the same. If there is a shortage of a particular part, that part limits the number of solar model cars that can be built. It is the limiting part.

To gain a better understanding of limiting reactants, let's look at a system that is not a chemical reaction. Suppose Joel wants to build a model solar-powered car. He needs 1 frame, 1 solar cell, 1 electric motor, and 4 wheels (Figure 6.8A). We can write an equation to represent the numbers of parts required to make the car:

1 frame + 1 solar cell + 1 electric motor + 4 wheels = 1 model solar car

To make one or more solar cars, the ratio of cells to motors to frames to wheels is 1:1:1:4. If the numbers of parts are present in the same ratio, then all the parts can be used to make cars. For example, if Joel has 10 cells, 10 electric motors, 10 frames, and 40 wheels, he can make 10 cars with no parts left over. If the number of parts is different from this ratio, then the number of cars will be limited by at least one of the parts, and Joel will have leftover parts. For example, suppose he has 5 solar cells, 6 motors, 4 frames, and 12 wheels, as shown in Figure 6.8B. How many cars can he make? He has enough frames to make 4 cars. He has more than enough cells and motors, but does he have enough wheels? No, he would need 16 wheels to go with the 4 frames, and he has only 12. So, the number of wheels limits the number of cars to 3. The wheels are the *limiting part*. He has too many of the other parts, so they are in excess (Figure 6.8C). Notice that the terms we are using are similar to those we use in discussing a chemical reaction. The difference is that we are describing parts instead of reactants.

EXAMPLE 6.4 | Limiting Parts

Suppose Joel received a shipment of 88 wheels, so he now has a total of 100. He still has 5 solar cells, 6 motors, and 4 frames.

(a) Use the parts ratio described in Figure 6.8A to determine the limiting part(s).
(b) How many cars can Joel make?
(c) Identify which parts are in excess and how many of each will be left over.

Solution:

(a) The large shipment of wheels provides excess wheels. They are no longer the limiting part. The numbers of cells, motors, and frames needed per car are the same (1:1:1). Since Joel has only 4 frames, and more of each of the other parts, the frame is the limiting part.
(b) The number of limiting parts (frames) determines the number of cars that can be built. Since each car requires 1 frame, a maximum of 4 cars can be built.
(c) Joel has 1 excess cell, 2 excess motors, and 84 excess wheels. They will be left over after 4 complete cars are constructed.

Practice Problem 6.4

Suppose Joel then got a shipment of 10 additional frames.

(a) Use the parts ratio described in Figure 6.8A to determine the limiting part(s).
(b) How many cars can Joel make?
(c) Identify which parts are in excess and how many of each will be left over.

Further Practice: Questions 6.31 and 6.32 at the end of the chapter

Now let's look at what occurs at the molecular level in a reaction that has a limiting reactant. Consider the combustion of hydrogen to form water. The molecular-level diagram and balanced equation in Figure 6.9 show that the molecule ratio of H_2 to O_2 is $\frac{2\ H_2\ \text{molecules}}{1\ O_2\ \text{molecule}}$. The *molecule ratio* is the same as the *mole ratio* because both are related to the coefficients in the balanced equation.

In this chemical reaction, the reactant molecules are the "parts" that combine in a specific ratio to make the product H_2O molecules. If the number of H_2 molecules mixed is exactly two times the number of O_2 molecules, then all the H_2 and O_2 molecules will react, and no reactants will be left over. If the ratio of H_2 to O_2 molecules mixed together is *not* exactly $\frac{2\ H_2\ \text{molecules}}{1\ O_2\ \text{molecule}}$, then one reactant is the limiting

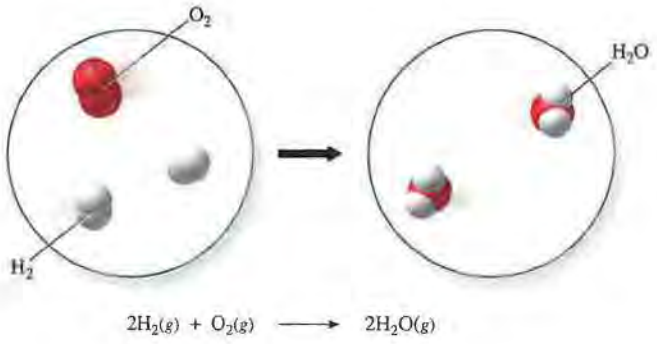


FIGURE 6.9 For every 2 H_2 molecules that react, 1 O_2 molecule reacts with it to form 2 H_2O molecules.

reactant and the other is in excess. For example, suppose 8 molecules of H_2 are mixed with 5 molecules of O_2, as shown in Figure 6.10. The reactant ratio is not $\frac{2\ H_2\ \text{molecules}}{1\ O_2\ \text{molecule}}$, so only one of the reactants is completely consumed. The other is in excess. After the reactants combine in this 2 to 1 ratio, 8 H_2O molecules form. The H_2 reactant is the limiting reactant because it was completely used up. It limited the number of H_2O molecules that could form. The O_2 reactant is left over. It was in excess. When the reaction is complete, the 1 leftover O_2 molecule that did not react will also be present in the reaction mixture along with the 8 H_2O molecules that formed.

Suppose there was no "after" picture in Figure 6.10. Could you still determine the limiting reactant? Yes, it could be done in a way similar to the method used to determine the limiting part for building model cars. The molecule ratio in the balanced equation shows the number of molecules of one reactant that is required to react with the given number of molecules of the other reactant. Let's arbitrarily start by determining the number of H_2 molecules needed to react with the 5 O_2 molecules given in Figure 6.10. We multiply the given number of O_2 molecules by the molecule ratio from the balanced equation to get the number of H_2 molecules:



$$\text{Molecules of } H_2 = 5\ \cancel{O_2\ \text{molecules}} \times \frac{2\ H_2\ \text{molecules}}{1\ \cancel{O_2\ \text{molecule}}} = 10\ H_2\ \text{molecules}$$

This calculation tells us that we need 10 H_2 molecules to react with the 5 O_2 molecules. Since we actually have only 8 H_2 molecules, we know that

$$\text{Calculated } H_2 > \text{actual } H_2$$

Since there is not enough H_2 to react with all the O_2, H_2 is the limiting reactant. We come to the same conclusion if we calculate the number of O_2 molecules required to react with the 8 H_2 molecules:

$$\text{Molecules of } O_2 = 8\ \cancel{H_2\ \text{molecules}} \times \frac{1\ O_2\ \text{molecule}}{2\ \cancel{H_2\ \text{molecules}}} = 4\ O_2\ \text{molecules}$$

This calculation tells us that we need 4 O_2 molecules to react with the 8 H_2 molecules. We actually have 5 O_2 molecules:

$$\text{Calculated } O_2 < \text{actual } O_2$$

We have 1 extra O_2 molecule, so O_2 is in excess. The H_2 must be the limiting reactant. All 8 H_2 molecules will react with 4 of the O_2 molecules, and there will be 1 leftover O_2 molecule. Does this match the "after" picture in Figure 6.10? It should. It makes no difference which reactant you start with to determine the limiting reactant.

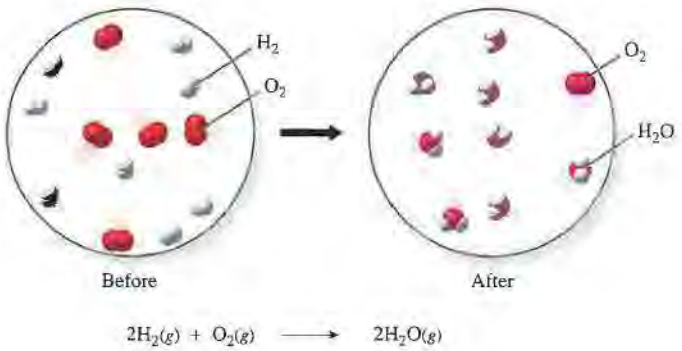


FIGURE 6.10 When 8 H_2 molecules and 5 O_2 molecules are combined, all of the H_2 reacts and 1 O_2 molecule is left over. The limiting reactant is H_2, and O_2 is in excess.

Your conclusion will be the same. The following table shows how much is mixed, how much reacts, and the final composition of the mixture.

	$2H_2(g)$	+	$O_2(g)$	⟶	$2H_2O(g)$
Initially mixed	8 molecules		5 molecules		0 molecules
How much reacts	8 molecules		4 molecules		
Composition of final mixture	0 molecules		1 molecule		8 molecules

The following summarizes the method for determining the limiting reactant.

Steps for Determining the Limiting Reactant

If you calculate the amount of product predicted from the given amount of each reactant, the calculation that gives the *least* amount of product identifies the limiting reactant.

1. Calculate the amount of one reactant (B) needed to react with the other reactant (A).
2. Compare the calculated amount of B (amount needed) to the actual amount of B that is present initially.
   (a) If calculated B = actual B, there is no limiting reactant. Both A and B will react completely.
   (b) If calculated B > actual B, B is the limiting reactant. Only B will react completely.
   (c) If calculated B < actual B, A is the limiting reactant. Only A will react completely.

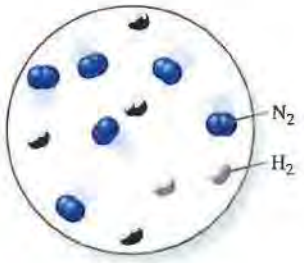
Example 6.5 will give you practice identifying limiting reactants in a chemical reaction.

EXAMPLE 6.5 | Limiting Reactants—Molecular Level

Consider the combination reaction of nitrogen gas with hydrogen gas to form ammonia gas:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

The molecular-level diagram shows the number of N_2 and H_2 molecules mixed together before the reaction occurs.



Before

(a) Identify the limiting reactant, assuming the reaction goes to completion.
(b) Draw the "after" picture. Include the correct number of product molecules and any leftover reactant molecules.
(c) Determine the reactant that is in excess.

Solution:

(a) The limiting reactant can be identified by determining the number of molecules of one reactant needed to react with the given number of molecules of

the other. Here we will calculate the number of N_2 molecules needed to react with the given number of H_2 molecules:

$$\text{Molecules of } N_2 = 6 \text{ } \cancel{H_2 \text{ molecules}} \times \frac{1 \text{ } N_2 \text{ molecule}}{3 \text{ } \cancel{H_2 \text{ molecules}}} = 2 \text{ } N_2 \text{ molecules}$$

MATH TOOLBOX 1.3

We need 2 N_2 molecules and we actually have 6.

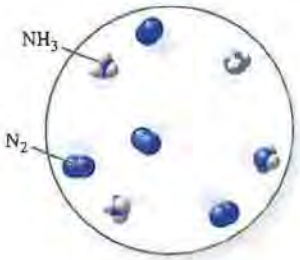
$$\text{Calculated } N_2 < \text{actual } N_2$$

We have 4 extra N_2 molecules, so N_2 is in excess. The H_2 must be the limiting reactant.

(b) Your "after" picture should have 4 NH_3 molecules and 4 N_2 molecules:

$$\text{Molecules of } NH_3 = 6 \text{ } \cancel{H_2 \text{ molecules}} \times \frac{2 \text{ } NH_3 \text{ molecules}}{3 \text{ } \cancel{H_2 \text{ molecules}}}$$

$$= 4 \text{ } NH_3 \text{ molecules}$$



After

(c) Because some N_2 molecules are left over, there was too much to react, so it is in excess. The following table summarizes this reaction.

	$N_2(g)$	+	$3H_2(g)$	⟶	$2NH_3(g)$
Initially mixed	6 molecules		6 molecules		0 molecules
How much reacts	2 molecules		6 molecules		
Composition of final mixture	4 molecules		0 molecules		4 molecules

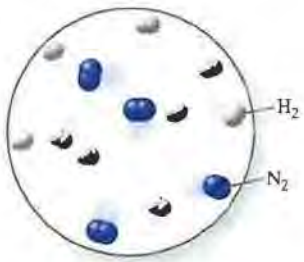
The last row in the table shows that H_2 reacted completely and some N_2 remains, so H_2 is the limiting reactant.

Practice Problem 6.5

Consider the combination reaction of nitrogen gas with hydrogen gas to form ammonia gas:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

The molecular diagram shows the number of N_2 and H_2 molecules mixed together before the reaction occurs.



Before

(a) Identify the limiting reactant, assuming the reaction goes to completion.
(b) Draw the "after" picture. Include the correct number of product molecules and any leftover reactant molecules.
(c) Determine the reactant that is in excess.

Further Practice: Questions 6.35 and 6.36 at the end of the chapter

Now that we have identified limiting reactants on a molecular scale, let's look at the much larger mole scale. The procedure is similar because the molecule ratio and the mole ratio are the same and are given by the balanced equation. Let's look at the reaction of sodium metal and chlorine gas:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Suppose we mix 0.50 mol Na with 0.20 mol Cl_2. What is the limiting reactant? We'll follow a similar procedure to what we did when we started with the number of molecules of each reactant. Let's begin by finding the moles of Na needed to react with the 0.20 mol Cl_2:



$$\text{Moles of Na} = 0.20 \cancel{\text{ mol Cl}_2} \times \frac{2 \text{ mol Na}}{1 \cancel{\text{ mol Cl}_2}} = 0.40 \text{ mol Na}$$

We need 0.40 mol Na, and we actually have 0.50 mol Na:

$$\text{Calculated Na} < \text{actual Na}$$

Na is in excess by 0.10 mol. The Cl_2 must be the limiting reactant. We can expect all the Cl_2 gas to react and some (0.10 mol) of the Na metal to be left over after the reaction.

We can determine how much NaCl should form by doing a mole-mole calculation. Which reactant do we use? We use the limiting reactant Cl_2 because it reacts completely. We multiply the moles of limiting reactant Cl_2 by the mole ratio of NaCl to Cl_2 to determine the moles of NaCl that can form:

$$\text{Moles of NaCl} = 0.20 \cancel{\text{ mol Cl}_2} \times \frac{2 \text{ mol NaCl}}{1 \cancel{\text{ mol Cl}_2}} = 0.40 \text{ mol NaCl}$$

The following table summarizes this reaction.

	$2Na(s)$	+	$Cl_2(g)$	$\longrightarrow$	$2NaCl(s)$
Initially mixed	0.50 mol		0.20 mol		0.00 mol
How much reacts	0.40 mol		0.20 mol		
Composition of final mixture	0.10 mol		0.00 mol		0.40 mol

When 0.50 mol Na is mixed with 0.20 mol Cl_2, 0.40 mol Na reacts with 0.20 mol Cl_2. When the reaction is complete, 0.40 mol NaCl is produced. Because the Na is in excess by 0.10 mol, that amount, along with the 0.40 mol NaCl that forms, will be present when the reaction is complete.

Example 6.6 will give you practice determining the limiting reactant and the amount of product that can form.

EXAMPLE 6.6 | Limiting Reactants—Mole Scale

The ethanol used in Lily's car was probably made from corn, but it also could be made from the hydrocarbon ethylene, C_2H_4. Ethylene, a gas at room temperature, burns with oxygen. The balanced equation is

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

Suppose we mix 0.25 mol C_2H_4 with 1.0 mol O_2. Identify the limiting reactant and the number of moles of CO_2 that can be formed.

Solution:

To determine the limiting reactant, first determine if there is enough C_2H_4 to react with 1.0 mol O_2:

$$\text{Moles of } C_2H_4 = 1.0 \cancel{\text{ mol } O_2} \times \frac{1 \text{ mol } C_2H_4}{3 \cancel{\text{ mol } O_2}} = 0.33 \text{ mol } C_2H_4$$



We need 0.33 mol C_2H_4, and we actually have only 0.25 mol C_2H_4:

$$\text{Calculated } C_2H_4 > \text{actual } C_2H_4$$

Because there is not enough C_2H_4 to react with all the O_2, C_2H_4 is the limiting reactant. (Assure yourself that you would reach the same conclusion if you started the calculation with the given amount of C_2H_4 instead of the given amount of O_2.) We use the amount of limiting reactant to calculate the moles of CO_2 product:

$$\text{Moles of } CO_2 = 0.25 \cancel{\text{ mol } C_2H_4} \times \frac{2 \text{ mol } CO_2}{1 \cancel{\text{ mol } C_2H_4}} = 0.50 \text{ mol } CO_2$$

We also use the amount of the limiting reactant to determine the amount of O_2 that will react:

$$\text{Moles of } O_2 \text{ that react} = 0.25 \cancel{\text{ mol } C_2H_4} \times \frac{3 \text{ mol } O_2}{1 \cancel{\text{ mol } C_2H_4}} = 0.75 \text{ mol } O_2$$

This reaction is summarized in the following table.

	$C_2H_4(g)$	+	$3O_2(g)$	⟶	$2CO_2(g)$	+	$2H_2O(g)$
Initially mixed	0.25 mol		1.00 mol		0.00 mol		0.00 mol
How much reacts	0.25 mol		0.75 mol				
Composition of final mixture	0.00 mol		0.25 mol		0.50 mol		0.50 mol

Practice Problem 6.6

In the double-displacement reaction of aqueous silver nitrate and magnesium chloride, silver chloride precipitates and aqueous magnesium nitrate forms:

$$2AgNO_3(aq) + MgCl_2(aq) \longrightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$$

Suppose we mix 1.0 mol of silver nitrate with 1.0 mol of magnesium chloride. Identify the limiting reactant and the number of moles of solid AgCl that can form.

Further Practice: Questions 6.41 and 6.42 at the end of the chapter

Generally we measure amounts of reactants in units of grams. Given grams of reactants, we must first convert from grams to moles for each reactant. Example 6.7 shows how to determine the limiting reactant when grams of reactants are given.



MATH TOOLBOX 1.3

EXAMPLE 6.7 | Limiting Reactants When Given the Mass of Reactants

When liquid bromine is mixed with aluminum metal, a combination reaction occurs, forming aluminum bromide:

$$2Al(s) + 3Br_2(l) \longrightarrow 2AlBr_3(s)$$

Determine the limiting reactant and mass of $AlBr_3$ produced when 45.0 g of Br_2 are added to 30.0 g of Al.

Solution:

Because we are given the mass of reactants, we must first convert from grams to moles for each reactant:

$$\text{Moles of Al} = 30.0 \cancel{\text{ g Al}} \times \frac{1 \text{ mol Al}}{26.98 \cancel{\text{ g Al}}} = 1.11 \text{ mol Al}$$

$$\text{Moles of } Br_2 = 45.0 \cancel{\text{ g } Br_2} \times \frac{1 \text{ mol } Br_2}{159.89 \cancel{\text{ g } Br_2}} = 0.281 \text{ mol } Br_2$$

Next we determine the number of moles of one reactant needed to react with the other. Here we will determine the moles of Br_2 needed to react with 1.11 mol Al:

$$\text{Moles of } Br_2 = 1.11 \cancel{\text{ mol Al}} \times \frac{3 \text{ mol } Br_2}{2 \cancel{\text{ mol Al}}} = 1.67 \text{ mol } Br_2$$

We need 1.67 mol Br_2, but we actually have only 0.281 mol Br_2:

$$\text{Calculated } Br_2 > \text{actual } Br_2$$

Because Br_2 is in short supply, it is the limiting reactant. (Convince yourself that you would come to the same conclusion if you started the calculation with the given amount of Br_2 instead of the given amount of Al.)

You may have been able to see merely by looking at the balanced equation that Br_2 is the limiting reactant. The balanced equation shows that the number of moles of Br_2 must be greater than that of Al, but the number of moles of Br_2 we have is less than that of Al. Also note that although the mass ratio of Br_2 to Al mixed together is (45 g)/(30 g), or 3/2, this is not the mole ratio. *When given the mass of reactants, we must always first convert to moles.*

To determine the mass of product, we can do a mass-mass conversion starting with the given grams of Br_2. Or, we can start with the moles of Br_2 calculated from the grams of Br_2, convert moles of Br_2 to moles of $AlBr_3$, and then convert moles of $AlBr_3$ to grams of $AlBr_3$:



$$\text{Mass of } AlBr_3 = 0.281 \cancel{\text{ mol } Br_2} \times \frac{2 \cancel{\text{ mol AlBr}_3}}{3 \cancel{\text{ mol } Br_2}} \times \frac{266.7 \text{ g AlBr}_3}{1 \cancel{\text{ mol AlBr}_3}}$$

$$= 50.0 \text{ g AlBr}_3$$

The maximum amount of $AlBr_3$ that can be produced is 50.0 g.

Practice Problem 6.7

Sodium carbonate reacts with hydrochloric acid to produce aqueous sodium chloride, water, and carbon dioxide gas. The balanced equation is

$$Na_2CO_3(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Determine the limiting reactant and mass of CO_2 gas produced when 11.0 g of Na_2CO_3 is added to a solution that contains 11.0 g of HCl.



Further Practice: Questions 6.43 and 6.44 at the end of the chapter

6.5 | PERCENT YIELD

When the manufacturers of E-85 produce ethanol to power Lily's car, they calculate the amount of ethanol product that *should* form from a reaction. This is the **theoretical yield**, the maximum amount of product that can be obtained from the given amounts of reactants. In Example 6.7 for instance, we calculated the theoretical yield of $AlBr_3$ as 50.0 g.

When chemists do experiments and when companies manufacture fuels, they seldom obtain the amount of product predicted by the calculation. They fall short of the theoretical yield for any number of reasons. Reactants and products can be lost due to spillage or during transfer from one container to another (Figure 6.11); and alternative reactions (side reactions) can occur that produce different products. The amount of product we measure in the laboratory is the **actual yield**. The actual yield is usually less than the theoretical yield; however, it can appear greater if a solid is weighed when still wet or if a product is contaminated.

The **percent yield** describes how much of a product is actually formed in comparison to how much should have been formed. It is the actual yield expressed as a percentage of the theoretical yield:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

The actual yield and theoretical yield can be expressed in any units, as long as those units are the same. Usually yields are reported in mass units, but yields in units of moles or molecules are acceptable.

Suppose the reaction of aluminum metal and liquid bromine described in Example 6.7 produced only 26.8 g $AlBr_3$. What is the percent yield? Since we calculated the amount of $AlBr_3$ that should form as 50.0 g, we know the theoretical yield.



FIGURE 6.11 The actual yield is usually less than the theoretical yield due to loss during reaction steps, purification, and transfer processes, or due to side reactions.

The percent yield is the ratio of the actual yield to the theoretical yield multiplied by 100:

$$\text{Percent yield} = \frac{26.8\text{ g}}{50.0\text{ g}} \times 100\% = 53.6\%$$

Example 6.8 will help you better understand percent yield, theoretical yield, and actual yield.

EXAMPLE 6.8 | Percent Yield, Theoretical Yield, and Actual Yield

Sodium metal reacts with water in a single-displacement reaction to produce aqueous sodium hydroxide and hydrogen gas. The balanced equation is

$$2\text{Na}(s) + 2\text{H}_2\text{O}(l) \longrightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$$

When 0.50 mol Na is placed in water, all the sodium metal reacts, and the hydrogen gas produced is isolated. It is determined that 0.21 mol H_2 has been produced. What is the percent yield of H_2?

Solution:

To calculate percent yield, we divide the actual yield by the theoretical yield, and then multiply by 100:

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

The actual yield is the number of moles of H_2 that was isolated once the reaction occurred (0.21 mol). The theoretical yield must be calculated from the amount of limiting reactant that reacted. In this case we can assume the Na is the limiting reactant (not the water) because all the Na reacted. We calculate the theoretical yield by calculating the maximum number of moles of H_2 that should be produced from the moles of limiting reactant Na:



$$\text{Moles of H}_2 = 0.50\ \cancel{\text{mol Na}} \times \frac{1\text{ mol H}_2}{2\ \cancel{\text{mol Na}}} = 0.25\text{ mol H}_2\text{ (theoretical yield)}$$

We substitute the values of actual and theoretical yield into the percent yield equation and solve for percent yield:

$$\text{Percent yield} = \frac{0.21\text{ mol H}_2}{0.25\text{ mol H}_2} \times 100\% = 84\%$$

Practice Problem 6.8

Toxic carbon monoxide is a by-product of the preparation of hydrogen fuel from methanol for use in hydrogen-powered vehicles. When carbon monoxide reacts with oxygen gas, the greenhouse gas carbon dioxide is produced:

$$2\text{CO}(g) + \text{O}_2(g) \longrightarrow 2\text{CO}_2(g)$$

When 5.0 mol CO is mixed with excess O_2, the reaction occurs to give 3.4 mol CO_2. What is the percent yield of CO_2?

Further Practice: Questions 6.55 and 6.56 at the end of the chapter

If we know the expected percent yield for a reaction, we can predict the amount of product that can realistically be made. Such calculations are important in industry where percent yield is a significant factor in determining profits.

6.6 | ENERGY CHANGES

When hydrogen and oxygen react to produce water, energy is released. In a hydrogen fuel cell, most of this energy is converted to electric energy, which can be used to run a vehicle. Suppose we combine the same amounts of hydrogen and oxygen, but without the fuel cell. Would the same amount of energy be released by the reaction? The answer is yes, but the energy would not be in such a usable form. Some would be lost as *heat* to the surroundings, causing the temperature of the environment to increase. This is what happens when we burn other fuels, such as propane in a gas grill or wood in a bonfire. Whether the energy is converted to useful work or is dissipated into the environment as heat, the amount of energy released is the same. So far in this chapter, we have looked only at the amounts of reactants and products in a chemical reaction. Here we will investigate the amount of energy absorbed or released by a chemical reaction.

See Section 1.3 for an introduction to energy.

Law of Conservation of Energy

As described in Chapter 1, energy can take many forms, and it can be converted from one form to another. Energy can also be transferred as heat. In either case, the amount of energy remains the same. The **law of conservation of energy** expresses this principle: Energy can be converted or transferred, but it cannot be created or destroyed. Thus, when we "use" energy, we are not using it up. We are converting it to a different form of energy. However, we may be converting it into forms that are not useful, such as heat lost to the surrounding environment. When gasoline is burned in an engine, the chemical energy in the gasoline is converted to mechanical energy and heat energy (Figure 6.12). The mechanical energy does useful work, but the heat energy is generally lost to the atmosphere. If gasoline is burned in a barrel, most of the energy would be lost as heat. In order to make the energy possessed by an object most useful to us, we try to minimize the amount of energy lost as heat.

Efficiency is a measure of the amount of useful work that is achieved from an energy conversion. No conversion is totally efficient, as shown in Figure 6.13. The process of conversion always generates some heat that is lost even before useful work can be done. The efficiency of a conversion is the percentage that ends up in the form we want. Consider, for example, the series of conversions from the chemical energy stored in coal to the light energy we see when we turn on a desk lamp (Figure 6.14). When coal is burned in a steam boiler, about 88% of its chemical energy is converted to heat energy, which can be used to heat water and form steam. A steam turbine at the power plant can convert about 46% of the heat energy in the steam to mechanical energy, and an electric generator can convert about 98% of the mechanical energy to electric energy. Finally, when the electric energy is used to operate a lightbulb, only about 4% is converted to light, while the rest is lost to the surroundings as heat. In the overall process of burning coal to operate a lightbulb,

Regular incandescent lightbulbs emit a large amount of energy as heat. Fluorescent lightbulbs generate light by a different process, which does not produce much heat. As a result, fluorescent lightbulbs are four to six times more efficient than incandescent lightbulbs. A 15-W fluorescent lightbulb delivers as much light as a 60-W incandescent lightbulb.

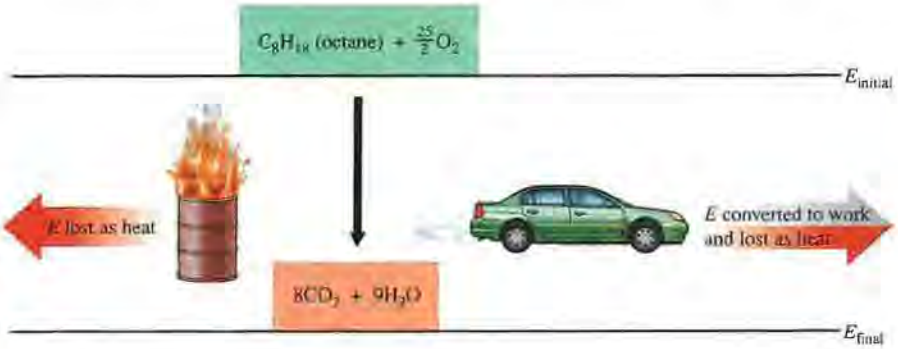


FIGURE 6.12 When gasoline burns in an ICE, about 32% of the energy released is converted to work and 68% is immediately lost to the surrounding environment as heat. When gasoline is burned in a barrel, 100% of the energy released is immediately lost as heat. The total energy released is the same.

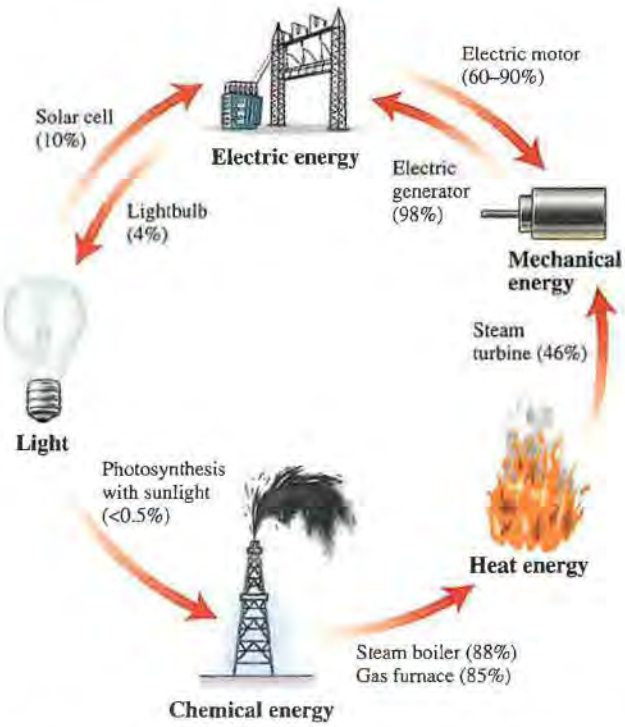


FIGURE 6.13 The conversion of one kind of energy to another is usually not completely efficient, because some energy is converted into undesired forms. Some typical conversion efficiencies are shown.

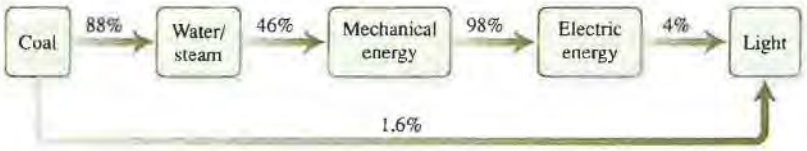


FIGURE 6.14 From power plant to desk lamp, the overall conversion efficiency is only 1.6%. The other 98.4% of coal's energy is converted to heat energy before ever performing the desired function.

only 1.6% of the chemical energy ends up as light energy. The rest has been lost along the way because of inefficient conversions. Most of the "lost" energy ends up as heat dissipated to the air or into the bodies of water used to cool power plants.

Energy Changes That Accompany Chemical Reactions

Lily wonders how the alternative fuel she uses provides the energy that moves her car from one place to another. Where does the energy come from? It comes from chemical energy that is stored as potential energy in the fuel. It is stored in the *bonds* that hold atoms together in molecules. Some substances release some of their potential energy when they react to form substances that store less potential energy. Ethanol, for example, stores a large amount of potential energy in the bonds that hold its atoms together. Some of that energy is released as heat when it reacts with oxygen. A reaction that releases energy is called an **exothermic reaction**.

Not all reactions release energy. Some require a continuous energy input. A reaction that absorbs energy from the surroundings is an **endothermic reaction**.

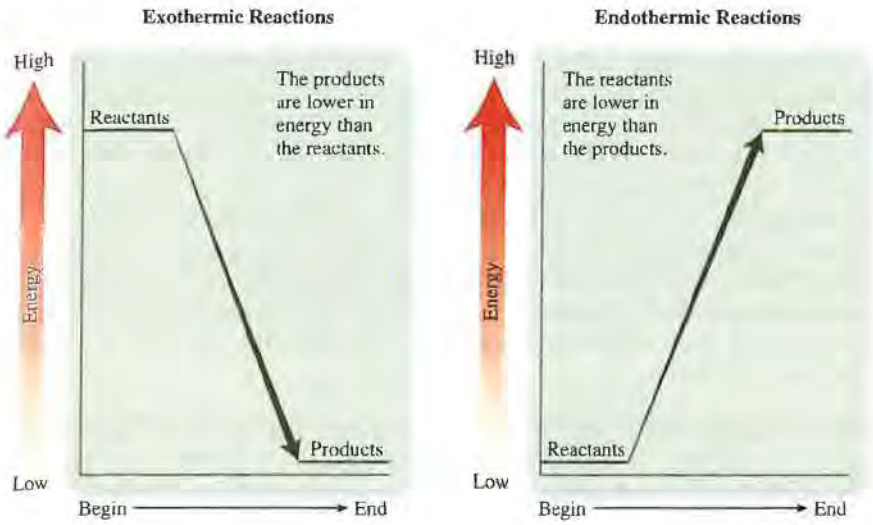


FIGURE 6.15 An exothermic reaction is thought of as being "downhill" because the products contain less stored energy than the reactants. An endothermic reaction is thought of as being "uphill" because the products contain more potential energy.

In an endothermic reaction, the products have more potential energy than the reactants. Figure 6.15 shows *reaction profiles,* which describe the overall energy changes that accompany exothermic and endothermic reactions. As shown by the endothermic reaction profile, the energy of the reactants is less than the energy of the products. The difference is the energy absorbed as the reaction takes place. In the exothermic reaction profile, the reactants contain more energy than the products. The difference is the energy released during the reaction. These reaction profiles also show us that if a reaction is exothermic, it will become endothermic if it is run in the reverse direction. And if a reaction is endothermic, its reverse reaction is exothermic.

Physical changes, such as phase changes, can also be endothermic or exothermic. Sweating is a natural process that helps cool our bodies when the water in sweat evaporates. Evaporation is an endothermic process, so it absorbs heat from our skin. Our skin cools, helping to maintain normal body temperature.

EXAMPLE 6.9 | Endothermic and Exothermic Reactions

The following is the reaction used in a hydrogen fuel cell:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

(a) Is this reaction endothermic or exothermic?
(b) The reverse of this reaction, the formation of hydrogen and oxygen from water, can be accomplished by running electricity through water. Is this reaction endothermic or exothermic? Explain.

Solution:

(a) The reaction releases the energy that is converted to useful energy by the fuel cell. Since energy is released by the reaction, it is exothermic.
(b) The reverse of this reaction requires a continuous energy input (electricity) to convert from reactants to products, so it is an endothermic reaction. If a reaction is exothermic, its reverse reaction is endothermic.

Practice Problem 6.9

When a car battery does not work, it can usually be recharged so that it is useful again. Recharging a battery involves running the reaction in reverse. Is this reaction an endothermic or exothermic process? Explain.

Further Practice: Questions 6.67 and 6.68 at the end of the chapter

Nutrition Facts

Serving Size: 1 cup (54g/1.9 oz.)
Servings Per Container: About 9

Amount Per Serving

Calories 190	**Calories from Fat** 10
	% Daily Value**
Total Fat 1g*	2%
Saturated Fat 0g	0%
Trans Fat 0g	
Cholesterol 0mg	0%
Sodium 0mg	0%
Potassium 180mg	5%
Total Carbohydrate 45g	15%
Dietary Fiber 6g	24%
Soluble Fiber 1g	
Insoluble Fiber 5g	
Sugars 7g	
Other Carbohydrates 32g	
Protein 5g	
Vitamin A 0% •	Vitamin C 0%
Calcium 0% •	Iron 8%

* Amount in cereal. One half cup of fat free milk contributes an additional 40 calories, 65mg sodium, 6g total carbohydrate (6g sugars), and 4g protein.

** Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.

	Calories:	2,000	2,500
Total Fat	Less Than	65g	80g
Sat. Fat	Less Than	20g	25g
Cholesterol	Less Than	300mg	300mg
Sodium	Less Than	2,400mg	2,400mg
Potassium		3,500mg	3,500mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g
Protein		50g	65g

Calories per gram:
Fat 9 • Carbohydrate 4 • Protein 4

INGREDIENTS: Organic Whole Grain Wheat, Organic Evaporated Cane Juice, Natural Flavor.

FIGURE 6.16 A nutritional label gives the energy content in Calories (or kilocalories) per serving. Also listed are grams of fat, carbohydrates, and proteins. This label also lists the Calories from fat.

Quantities of Heat

For a chemical reaction, the energy change that is easiest to measure is the *heat* absorbed or released. In order to understand how we measure heat changes, we must first learn more about heat and how it is related to temperature changes. **Heat** is energy that is transferred between two objects because of a difference in their temperatures. First think about the heat associated with some physical changes. If we place warm pancakes on a plate, for example, heat is transferred from the pancakes to the plate. The pancakes lose heat and get cooler, while the plate gains heat and gets hotter until the temperature of the pancakes and the temperature of the plate are the same. Heat is always transferred from the hotter to the colder object. When two objects that are in contact become equal in temperature, they have reached a state of *thermal equilibrium.*

Units of Energy Many packaged foods are labeled with the energy content. For example, a small box of raisins contains 130 Calories (Cal with a capital *C*). How much energy is this? Nutritionists and chemists use related, but not identical, units to measure energy. Chemists measure energy in units of *joules* (J) or *calories* (cal with a lowercase *c*). One calorie is the amount of heat energy needed to raise the temperature of 1 g of water by 1°C. A joule is smaller than a calorie:

$$4.184\ \text{J} = 1\ \text{cal}$$

A *kilojoule* (kJ), or 1000 J, is approximately the amount of energy that is emitted when a wooden kitchen match burns completely. The calorie used in chemistry should not be confused with the *Calorie* used by nutritionists (Figure 6.16), which is actually a kilocalorie (kcal), or 1000 cal.

EXAMPLE 6.10 | Units of Energy

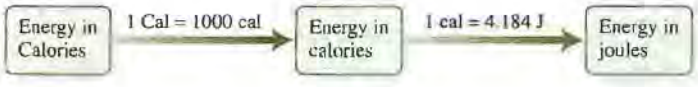
A can of cola from the United States contains 180 Cal, and a can containing the same amount of cola from Australia contains 900 J. Which contains more energy? Which is the diet cola?

Solution:

To answer this question, we must compare the energy in the two drinks using the same units. Thus, we must convert one value into the same unit as the other value. We will convert energy in Calories to energy in joules:



We have no direct relationship between these two quantities, so we have to break this down into two steps. First, note that the relationships between these two quantities are 1 Cal = 1000 cal and 1 cal = 4.184 J:



$$\text{Energy in J} = 180\ \cancel{\text{Cal}} \times \frac{1000\ \cancel{\text{cal}}}{1\ \cancel{\text{Cal}}} \times \frac{4.184\ \text{J}}{1\ \cancel{\text{cal}}} = 7.5 \times 10^5\ \text{J}$$

MATH TOOLBOX 1.3

This is considerably more energy than the 900 J found in the Australian cola, which must be the diet cola. Another way to solve the problem would be to convert 900 J to Calories and make the comparison.

Practice Problem 6.10

How many joules of energy are in a candy bar that contains 235 Cal?

Further Practice: Questions 6.79 and 6.80 at the end of the chapter

Specific Heat When heat is added to a substance, the temperature of that substance increases. The amount of heat that must be added to 1 g of a substance to raise its temperature by 1°C is a unique property of that substance, called its **specific heat.** Specific heat values are expressed in units of joules per gram per degree Celsius [J/(g °C)] or calories per gram per degree Celsius [cal/(g °C)]. The specific heat of liquid water is 1.000 cal/(g °C) or 4.184 J/(g °C). This means that if we have a 1.00-g sample of water and we want to increase its temperature by 1.00°C, we must add 1.00 cal or 4.18 J of heat. Similarly, if we have a 2.00-g sample of water and want to increase its temperature by 1.00°C, we must add 2.00 cal or 8.37 J of heat. The more of that substance we have, the more heat we need to add. How much energy will increase the temperature of a 1.00-g sample by 10.0°C? You need 10.0 times as much heat for a 10.0°C increase than for a 1.00°C increase; you need 10.0 cal or 41.8 J. Some representative values of specific heat for various substances are presented in Table 6.2.

When you eat pizza fresh from the oven, why do you burn the top of your mouth and not your tongue? One factor is that the specific heat of cheese is greater than that of the crust. The cheese transfers a greater amount of heat for a given temperature change. Another factor is the greater ability of the cheese to conduct heat. It transfers heat faster.

TABLE 6.2 | Specific Heats of Some Substances

	Specific Heat			Specific Heat	
Substance	**J/(g °C)**	**cal/(g °C)**	**Substance**	**J/(g °C)**	**cal/(g °C)**
aluminum(*s*)	0.895	0.214	water(*s*)	2.027	0.484
carbon (diamond)	0.508	0.121	water(*l*)	4.184	1.000
carbon (graphite)	0.708	0.169	water(*g*)	2.015	0.482
calcium(*s*)	0.656	0.157	asphalt	0.920	0.220
chromium(*s*)	0.450	0.108	bone	0.440	0.105
copper(*s*)	0.377	0.0900	brick	0.84	0.20
gold(*s*)	0.129	0.0310	cheddar cheese	2.60	0.621
iodine(*s*)	0.214	0.0510	concrete	0.88	0.21
iron(*s*)	0.448	0.107	glass	0.84	0.20
lead(*s*)	0.129	0.0310	granite	0.79	0.19
mercury(*l*)	0.140	0.0335	marble	0.86	0.21
silver(*s*)	0.234	0.0560	olive oil	1.79	0.428
tin(*s*)	0.222	0.0530	sand	0.835	0.200
uranium(*s*)	0.117	0.0280	strawberries	3.89	0.930
			wax	2.89	0.69

The amount of heat absorbed by a substance when it increases in temperature equals the amount of heat released when it cools back down to the same temperature. The amount of heat transferred to or from a substance is related to the mass, the specific heat, and the change, or difference, in temperature.

EXAMPLE 6.11 | Factors Affecting Heat

Samples of different metals having the same mass are heated to the same temperature and are then allowed to cool in wax. When all have cooled to the same temperature, they have melted through different amounts of the wax. Which of the metals has the greatest specific heat?



Solution:

From the figure, we see that metal B melted the wax to the lowest level. It did so because it released the greatest amount of heat. Consider the factors that determine the amount of heat transferred. Greater heat could arise from a greater mass, greater specific heat, or greater temperature change. We started with samples of the same mass, and they all underwent the same temperature change, so these are not factors. Therefore, the specific heat must be greater for metal B than for the other metals.

Practice Problem 6.11

A drop of boiling water spilled on your hand stings, but a cup of boiling water spilled on your hand burns. What factor explains the difference?

Further Practice: Questions 6.81 and 6.82 at the end of the chapter

If we know the specific heat of an object, its mass, and its change in temperature, we can determine the amount of heat absorbed or released. We can calculate the heat (q) associated with a temperature change by multiplying the mass in grams m of a substance by its specific heat C and by the temperature change ΔT (which is the difference between the final temperature and the initial temperature, $T_f - T_i$):

$$q = m \times C \times \Delta T$$

Heat — Mass — Specific heat — Temperature change

When the temperature of a substance increases (ΔT is positive), q has a positive value; heat is absorbed by the substance and it gets hotter. When the temperature decreases (ΔT is negative), q is negative; heat is released by the substance and it cools down.

We can't measure heat directly. There is no heat-measuring device. Instead, we measure the initial and final temperatures of an object and calculate heat from the specific heat, mass, and temperature change. Example 6.12 demonstrates this.

EXAMPLE 6.12 | Heat and Temperature Change

A simple solar energy heater is a glass-topped box containing rocks. Sunlight heats the rocks during the day, and air is blown over them at night to heat a house. The box contains 7.5×10^4 g of rocks, which have a specific heat of 0.49 J/(g °C). How much additional heat can the rocks store when they warm from a nighttime temperature of 18°C (65°F) to a daytime temperature of 43°C (110°F)?

More sophisticated solar energy systems use silicon semiconductor panels that convert sunlight into electricity.



Solution:

The temperature of the rocks increases, so they absorb heat, and we expect the calculated value of q to be positive. To calculate heat we use the following equation:

$$q = m \times C \times \Delta T$$

The mass m and specific heat C values are given. The temperature change ΔT is not given, but we can calculate ΔT from the final and initial temperatures:

$$\Delta T = T_f - T_i = 43°C - 18°C = 25°C$$

Substituting values of mass, specific heat, and temperature change into the equation gives the heat change for the rocks:

$$\begin{aligned} q &= m \times C \times \Delta T \\ &= 7.5 \times 10^4 \cancel{g} \times 0.49 \frac{\text{J}}{\cancel{g}\,\cancel{°C}} \times 25 \cancel{°C} \\ &= 7.5 \times 10^4 \times 0.49\ \text{J} \times 25 \\ &= 9.2 \times 10^5\ \text{J} \quad \text{or} \quad 9.2 \times 10^2\ \text{kJ} \end{aligned}$$

Note that q is positive as we predicted, which means that the rocks have absorbed heat.

Practice Problem 6.12

The specific heat of aluminum is 0.895 J/(g °C). If 156 g of aluminum at 75.0°C is cooled to 25.5°C, how much heat is transferred? What is the sign of q, and what is its significance?

Further Practice: Questions 6.85 and 6.86 at the end of the chapter

We have calculated heat from known values of mass, specific heat, and temperature change. Since the heat equation uses just these four quantities, any one can be determined if the other three are known.

In the examples we have looked at so far, we have focused on the amount of heat absorbed or released by an object such as a rock or a metal. Sometimes, however, we are interested not in an object, but in a process such as a chemical reaction. We call whatever we are interested in—whether an object or a

A *system* can be an object such as a piece of pipe, or a process, such as a physical or a chemical change. The *surroundings* is everything immediately around the system. The surroundings is commonly the air or the liquid around the system.

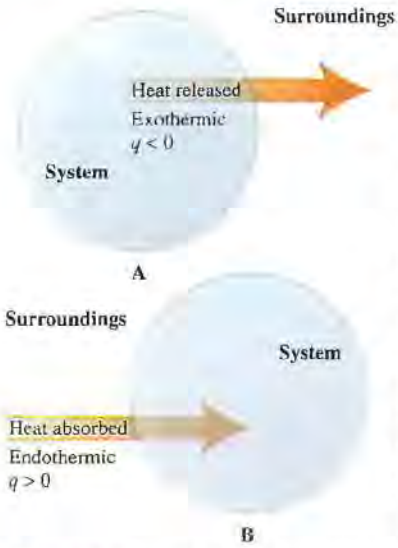


FIGURE 6.17 (A) When the system releases heat to the surroundings, q is a negative value ($q < 0$). (B) When the system absorbs heat from the surroundings, q is a positive value ($q > 0$).

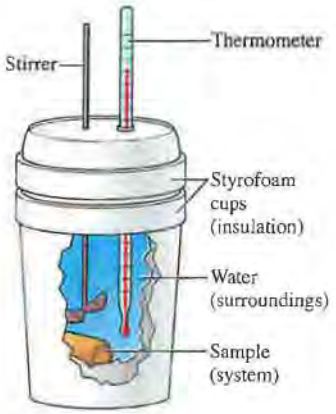


FIGURE 6.18 The amount of heat released by an object can be determined by measuring the amount of heat absorbed by the water, assuming the thermometer and stirrer absorb no heat. This experimental process is called *calorimetry*. The insulated container is called a *calorimeter*.

process—the *system*. When a system absorbs heat, where does the heat come from? When a system releases heat, where does the heat go? We know from the law of conservation of energy that energy is not created or destroyed. Heat absorbed by a system comes from its *surroundings*, which might be the air around it—or if in solution—the water around it. Heat released by a system is transferred to its surroundings.

Consider a car engine that heats up due to the combustion reaction in the cylinder. If we describe the system as the combustion reaction in the cylinder, is the value of q for the system positive or negative? We can't take the temperature of the combustion reaction, but we can tell that the surroundings of the combustion reaction (mostly the engine) increases in temperature. The car's engine must be absorbing heat. From the law of conservation of energy, we know that the heat is coming from the combustion reaction. The combustion reaction is releasing heat, so its value of q is negative. Figure 6.17 shows the relationship between a system, its surroundings, and heat transfer.

For some substances such as water, it's straightforward to calculate the amount of heat absorbed or released. We know the specific heat, and mass is easily measured. The initial and final temperatures of the water are easily measured, too. But the calculation is more difficult for an object such as a metal pipe. We might not know its specific heat, and we can't directly measure the temperature change of the pipe. Instead, we must design a heat transfer process that will let us determine the amount of heat transferred from the pipe to its surroundings, such as a known amount of water. In this case, we call the piece of pipe the system because it's what we are interested in. If we ensure that heat does not escape to anything but the water, then the water is the surroundings. There is no net loss of heat, so the heat change of the system plus the heat change of the surroundings equals zero:

$$q_{\text{system}} + q_{\text{surroundings}} = 0$$

In this example

$$q_{\text{pipe}} + q_{\text{water}} = 0$$

Suppose a 92.0-g piece of pipe is heated and then placed in an insulated container filled with 100.0 g of water at 25.00°C (Figure 6.18). The final temperature of the mixture is 29.45°C. If we assume no heat is lost from the water to the container, we can calculate the amount of heat the pipe loses to the water. It is the same as the amount of heat the water gains, which equals the mass of the water multiplied by the specific heat of water and its temperature change:

$$\begin{aligned} q_{\text{water}} &= m \times C \times \Delta T \\ &= 100.0 \cancel{\text{g}} \times 4.184 \frac{\text{J}}{\cancel{\text{g}}\ ^\circ\cancel{\text{C}}} \times (29.45\ \cancel{^\circ\text{C}} - 25.00\ \cancel{^\circ\text{C}}) \\ &= 1860\ \text{J} \end{aligned}$$

From the law of conservation of energy, the heat change of the pipe plus the heat change of the water equals zero:

$$q_{\text{pipe}} + q_{\text{water}} = 0$$

The heat change for the pipe, then, is the same value, but with the opposite sign:

$$q_{\text{pipe}} = -q_{\text{water}} = -1860\ \text{J}$$

The process just described is commonly used to determine the heat change for an object for which we do not know the specific heat, or if the initial temperature of the object is difficult to measure. The insulated container is commonly called a *calorimeter*. The measurement of heat transfer using a calorimeter is called *calorimetry*.

EXAMPLE 6.13 | Calorimetry

A brick is placed in an insulated calorimeter containing 5.000 kg of water. The temperature of the water decreases from 25.0°C to 19.4°C when thermal equilibrium is reached.

(a) Was the initial temperature of the brick greater than or less than the initial temperature of the water? Explain.
(b) What is the heat change q of the brick?

Solution:

(a) The temperature of the water decreased, so the temperature of the brick must have increased. Since they both have the same temperature at thermal equilibrium, the brick must have started out at a lower temperature than the initial temperature of the water.
(b) We don't know the mass or temperature of the brick, so we can't determine the heat change q for the brick directly. However, we do know the mass, specific heat, and temperature change of the water. The water is the brick's surroundings, so if we determine the heat change of the water, we can use the law of conservation of energy to determine the heat change of the brick. To determine the heat change of the water, we use the equation

$$q_{water} = m \times C \times \Delta T$$

The temperature change ΔT is the difference between the final temperature and the initial temperature:

$$\begin{aligned}\Delta T &= T_f - T_i \\ &= 19.4°C - 25.0°C \\ &= -5.6°C\end{aligned}$$

The mass of the water must be in units of grams, so we must convert kilograms to grams:

$$\text{Mass in grams} = 5.000 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 5.000 \times 10^3 \text{ g}$$

Substituting values for the mass of the water, specific heat of the water, and the temperature change of the water, we get

$$\begin{aligned}q_{water} &= 5.000 \times 10^3 \text{ g} \times 4.184 \frac{\text{J}}{\text{g °C}} \times (-5.6 \text{ °C}) \\ &= -1.2 \times 10^5 \text{ J}\end{aligned}$$

The negative value of q_{water} means that the water released heat. From the law of conservation of energy, we know that the brick absorbed the same amount of heat that the water released. The heat change of the brick plus the heat change of the water must sum to zero. This means that the heat change of the water is equal and opposite in sign to the heat change of the brick:

$$\begin{aligned}q_{brick} + q_{water} &= 0 \\ q_{brick} = -q_{water} &= +1.2 \times 10^5 \text{ J}\end{aligned}$$

Practice Problem 6.13

A sample of a metal alloy is heated and then placed in 125.0 g of water held in a calorimeter at 22.5°C. The final temperature of the water is 29.0°C. Assume heat exchange only occurs between the water and the alloy.

(a) Was the initial temperature of the alloy greater than or less than the initial temperature of the water? Explain.
(b) What is the heat change of the alloy?

Further Practice: Questions 6.89 and 6.90 at the end of the chapter

6.7 | HEAT CHANGES IN CHEMICAL REACTIONS

Lily wonders how the amount of heat released by combustion of a fuel such as ethanol is determined. A common way is to use a special type of calorimeter called a bomb calorimeter (Figure 6.19). It's similar to the calorimeter used for objects like pipes and bricks, but the reaction is kept separate from the water in a compartment that contains an ignition wire. A specific amount of the reactant is placed in the compartment. A small amount of electricity sent through the ignition wire gets the reaction going. Since all combustion reactions are exothermic, the reaction releases heat that the water absorbs.

By measuring the temperature change of the water, we can calculate the heat change of the water, q_{water}, using its specific heat. Assuming no heat is absorbed by the calorimeter, we can determine the heat change for the reaction, $q_{reaction}$, using the law of conservation of energy:

$$q_{reaction} + q_{water} = 0$$

$$q_{reaction} = -q_{water}$$

A bomb calorimeter can be used to determine the Calorie (energy) content of foods. Foods are fuels. They release their energy slowly when oxidized in the body or rapidly when burned in a calorimeter. Either way, the amount of energy is the same.

For example, if 5.0 g of a combustible substance burns in a calorimeter, and the surrounding water absorbs 48 kJ of heat ($q_{water} = +48$ kJ), then the combustion reaction must have released 48 kJ of heat:

$$q_{reaction} = -q_{water} = -48 \text{ kJ}$$

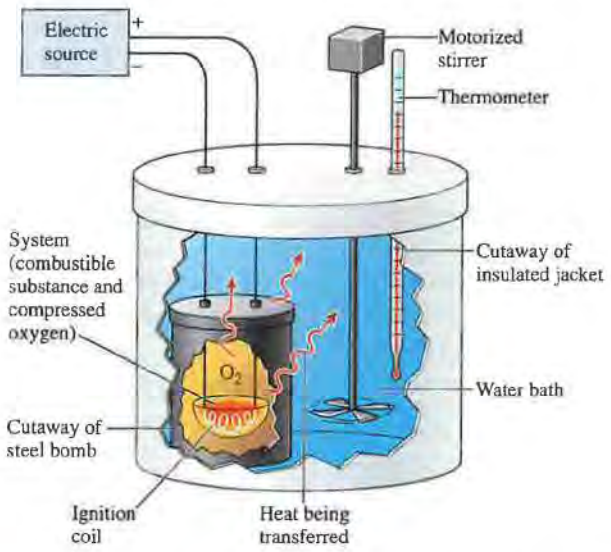


FIGURE 6.19 A bomb calorimeter is commonly used to measure the amount of heat released in combustion reactions. The water absorbs the heat released by the reaction.

The amount of energy released depends on the amount of substance burned. A few drops of a substance will produce much less heat than a gallon of that substance. In combustion reactions, it's common to report the heat change in terms of the amount of the substance that reacts with oxygen. One common way to report heat changes is in units of *kilojoules per gram* (kJ/g). In the previous example, 5.0 g of compound burned. The heat change *per gram* is one-fifth of the heat change when 5.0 g burns:

$$\frac{-48 \text{ kJ}}{5.0 \text{ g}} = -9.6 \text{ kJ/g}$$

Another way to report heat changes in chemical reactions is in units of *kilojoules per mole* (kJ/mol). Consider the combustion of propane gas, C_3H_8:

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

When 1 mol of propane burns, 2220 kJ of heat is released. The heat change for the combustion of propane, then, is –2220 kJ per mol, or

$$q_{\text{reaction}} = -2220 \text{ kJ/mol}$$

This ratio allows us to calculate the heat change when any molar amount of propane reacts. For example, if 2.00 mol of propane reacts, twice as much heat is released. This can be calculated by multiplying the 2.00 mol propane by the heat change per mole:



$$\text{Heat change} = 2.00 \cancel{\text{ mol } C_3H_8} \times \frac{-2220 \text{ kJ}}{1 \cancel{\text{ mol } C_3H_8}} = -4440 \text{ kJ}$$

EXAMPLE 6.14 | Heat Changes in Chemical Reactions

The ethanol, CH_3CH_2OH, in Lily's car burns with excess oxygen according to the following equation:

$$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

The heat change q for this reaction is –1367 kJ per mole of CH_3CH_2OH that reacts.

(a) Is this reaction endothermic or exothermic?
(b) What is the heat change when 0.200 mol of ethanol burns?

Solution:

(a) The value of q is negative, so the reaction is exothermic; energy is released by the reaction.
(b) To calculate the heat change when 0.200 mol CH_3CH_2OH reacts, multiply 0.200 mol CH_3CH_2OH by the heat change per mole of CH_3CH_2OH:

$$\text{Heat change} = 0.200 \cancel{\text{ mol } CH_3CH_2OH} \times \frac{-1367 \text{ kJ}}{1 \cancel{\text{ mol } CH_3CH_2OH}} = -273 \text{ kJ}$$

When 0.200 mol CH_3CH_2OH reacts, 273 kJ of heat is *released* to the surroundings.

Practice Problem 6.14

The balanced equation for the combination reaction of hydrogen gas and solid iodine is

$$H_2(g) + I_2(s) \longrightarrow 2HI(g)$$

The heat change q for this reaction is +53.00 kJ per mole of I_2 that reacts.

(a) Is this reaction endothermic or exothermic?
(b) What is the heat change when 2.50 mol of I_2 reacts?

Further Practice: Questions 6.99 and 6.100 at the end of the chapter

The combustion of ethanol releases less energy than the combustion of the hydrocarbons, like octane, that are found in gasoline. Thus, ethanol, although being used as a gasoline additive or replacement, is a less efficient fuel than gasoline. On the other hand, fuels derived from plant materials are renewable, consistent with the principles of green chemistry.

Foods and fuels are usually sold by volume or weight, so the kJ/g ratio is often useful. A mole of liquid octane, C_8H_{18}, for example, has a much greater mass and volume than a mole of liquid propane, C_3H_8. Table 6.3 shows the energy changes that correspond to the combustion of 1 g of a variety of fuels. Note that the energy released *per mole* generally increases with the number of carbon atoms in the molecular formula. However, the energy released *per gram* is relatively constant for compounds that contain only carbon and hydrogen. The energy released per gram is lower for methanol and ethanol, which also contain oxygen. Hydrogen provides the most energy per gram, which is one reason why it offers promise as a fuel of the future.

Fuels aren't the only substances that can be described in terms of energy content and mass. The energy content of foods is also reported in energy units per gram, and most of the world reports the energy content of foods in kilojoules. In the United States, the preferred unit is Cal/g, which is the same as kcal/g. The energy content of foods comes primarily from proteins, fats, and carbohydrates. The energy values of these components are listed in Table 6.4. Note that fats have a much higher energy value per gram than proteins or carbohydrates. This is one reason why many weight watchers limit the amount of fat in their diet.

Notice in Table 6.3 that a relatively small amount of energy is released when TNT (trinitrotoluene) burns. Other fuels release much more energy per mole or per gram. Is this surprising? The explosive nature of TNT has to do with its detonation *velocity*. The *rate* at which the reaction occurs is greater than for other fuels.

TABLE 6.3 | Heat Changes for the Combustion of Fuels
Substance + $xO_2(g) \longrightarrow yCO_2(g) + zH_2O(l)$

Fuel	Formula	Heat Change	
		kJ/mol	kJ/g
hydrogen	$H_2(g)$	−286	−143
methanol	$CH_3OH(l)$	−726	−23
methane	$CH_4(g)$	−891	−56
ethanol	$CH_3CH_2OH(l)$	−1367	−30
acetylene	$C_2H_2(g)$	−1301	−50
ethylene	$C_2H_4(g)$	−1411	−50
ethane	$C_2H_6(g)$	−1561	−52
propane	$C_3H_8(g)$	−2219	−50
butane	$C_4H_{10}(l)$	−2878	−50
octane	$C_8H_{18}(l)$	−5471	−48
wood	—	—	−15
TNT	$C_7H_5(NO_2)_3(s)$	−3434	−15

TABLE 6.4 | Energy Values of Food Components

Food Component	kcal/g (Cal/g)	kJ/g
proteins	4	17
fats	9	38
carbohydrates	4	17

The data in Table 6.4 are used to determine the energy content of foods as presented in most nutritional labels (Figure 6.16). Instead of burning every food in a calorimeter, food industry scientists determine the grams of protein, fat, and carbohydrate present in one serving. The mass of each is multiplied by their respective energy values. The calories from proteins, fats, and carbohydrates are then added to get the total energy content. A typical food label also lists the number of calories from fat.

SUMMARY

In this chapter we used balanced equations to determine quantities of reactants and products involved in chemical reactions. The coefficients in a balanced equation can be combined to give mole ratios (or molecule ratios), which are relationships between the moles of one reactant or product and another in the same reaction. If the mass of a reactant or product is given, the mass must first be converted to moles using the molar mass of the substance.

When a chemical reaction occurs, the reactants are usually not present in exactly the right amounts for a complete reaction of all reactants. In such cases, only the limiting reactant reacts completely. Once the limiting reactant is identified, it is used to determine the theoretical yield, or the maximum amount of product that can form. If the reaction is done in the laboratory, the amount of product actually obtained can be compared to the theoretical yield by calculating a percent yield.

Energy changes accompany chemical and physical changes. Endothermic reactions absorb energy from the surroundings; exothermic reactions release energy to the surroundings. Generally the energy transferred is in the form of heat. Temperature changes indicate the amount of heat transferred to or from an object. Specific heat is the amount of heat needed to raise the temperature of 1 g of a substance by 1°C. The relationship between the heat change of a substance, q, its mass m, specific heat C, and temperature change ΔT is given by the equation

$$q = m \times C \times \Delta T$$

Sometimes the amount of heat absorbed or released by a system—whether an object or a chemical reaction—cannot be determined directly. Instead, the heat change of the surroundings, usually water, is determined. As described by the law of conservation of energy, the heat change for the system is equal to the heat change of the surrounding water, but opposite in sign. Calorimetry is used to determine the heat change for many chemical reactions. The amount of heat absorbed or released during a chemical reaction depends on the mass of each reactant. The greater the amount of reactants, the greater the heat change. Heat changes for chemical reactions are usually reported in units of kJ/mol or kJ/g of a specific reactant.

KEY RELATIONSHIPS

Relationship	Equation
The percent yield for a reaction is a ratio of the amount of product actually obtained in a chemical reaction to the theoretical yield, the amount calculated from the limiting reactant.	$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$
The heat change for an object is equal to the product of its mass, its specific heat, and its temperature change. The temperature change ΔT is the difference between the final temperature and the initial temperature.	$q = m \times C \times \Delta T$

KEY TERMS

actual yield (6.5)
endothermic reaction (6.6)
exothermic reaction (6.6)
heat (6.6)
law of conservation of energy (6.6)
limiting reactant (6.4)
percent yield (6.5)
specific heat (6.6)
stoichiometry (6.1)
theoretical yield (6.5)

QUESTIONS AND PROBLEMS

The following questions and problems, except for those in the *Additional Questions* section are paired. Questions in a pair focus on the same concept. Answers to the odd-numbered questions and problems are in Appendix D.

Matching Definitions with Key Terms

6.1 Match the key terms with the descriptions provided.
(a) the process of determining the amounts of reactants and products in a chemical reaction
(b) the energy that is transferred between objects due to a difference in their temperatures
(c) a chemical change that absorbs heat
(d) the amount of heat required to raise the temperature of 1 g of a substance by 1°C
(e) the amount of product actually obtained in the laboratory from a reaction

6.2 Match the key terms with the descriptions provided.
(a) the reactant that is completely used up in a reaction and determines the amount of product that can form
(b) a chemical change that releases heat
(c) the maximum amount of product that can be obtained in a chemical reaction from known amounts of reactants; the amount of product calculated by assuming all of the limiting reactant is consumed
(d) the principle that states that the total amount of energy remains the same no matter how a physical or chemical change occurs
(e) the amount of a product actually obtained as a proportion of the amount expected, expressed as a percentage

The Meaning of a Balanced Equation

6.3 When one molecule of glucose, $C_6H_{12}O_6$, is metabolized in our bodies, it combines with six O_2 molecules to form six CO_2 molecules and six H_2O molecules.
(a) Write a balanced equation for this reaction.
(b) How many CO_2 molecules are formed when 12 glucose molecules react?
(c) If 30 O_2 molecules react, how many glucose molecules also react?

6.4 When two molecules of methanol, CH_3OH, react with oxygen, they combine with three O_2 molecules to form two CO_2 molecules and four H_2O molecules.
(a) Write a balanced equation for this reaction.
(b) How many H_2O molecules are formed when 10 methanol molecules react?
(c) If 30 O_2 molecules react, how many methanol molecules react?

6.5 What do the coefficients in a balanced equation represent?

6.6 Why is a balanced equation important when determining the amount of product that can be produced from a given amount of reactant?

6.7 The balanced equation for the dissolving of magnesium nitrate in water is

$$Mg(NO_3)_2(s) \xrightarrow{H_2O} Mg^{2+}(aq) + 2NO_3^-(aq)$$

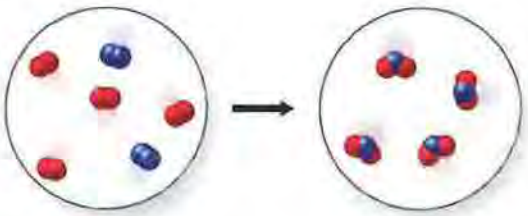
(a) How many $Mg^{2+}(aq)$ and $NO_3^-(aq)$ ions form for each $Mg(NO_3)_2$ formula unit that dissolves?
(b) How many moles of Mg^{2+} and NO_3^- form for each mole of $Mg(NO_3)_2$ that dissolves?

6.8 The balanced equation for the dissolving of sodium phosphate in water is

$$Na_3PO_4(s) \xrightarrow{H_2O} 3Na^+(aq) + PO_4^{3-}(aq)$$

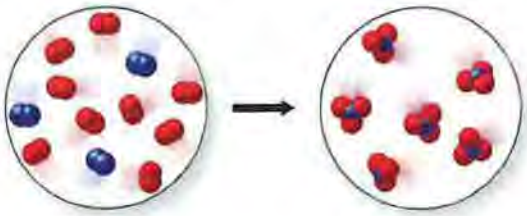
(a) How many Na^+ and PO_4^{3-} ions form for each Na_3PO_4 formula unit that dissolves?
(b) How many moles of Na^+ and PO_4^{3-} form for each mole of Na_3PO_4 that dissolves?

6.9 The molecular-level diagram shows the reactants (before the reaction) and the products (after the reaction). Which of the following is the best balanced equation to represent this reaction? For the choices you did not select, explain why they are incorrect.



(a) $2A + 4B \longrightarrow 4AB$
(b) $A_2 + B_2 \longrightarrow AB_2$
(c) $2A_2 + 4B_2 \longrightarrow 4AB_2$
(d) $A_2 + 2B_2 \longrightarrow 2AB_2$

6.10 The molecular-level diagram shows the reactants (before the reaction) and the products (after the reaction). Which of the following is the best balanced equation to represent this reaction on any scale? For the choices you did not select, explain why they are incorrect.



(a) $A_2 + B_2 \longrightarrow AB_3$
(b) $A_2 + 2B_2 \longrightarrow 2AB_3$
(c) $A_2 + 3B_2 \longrightarrow 2AB_3$
(d) $3A_2 + 9B_2 \longrightarrow 6AB_3$

Mole-Mole Conversions

6.11 Benzene, C_6H_6, is a known carcinogen that burns in air according to the following equation:

$$2C_6H_6(l) + 15O_2(g) \longrightarrow 12CO_2(g) + 6H_2O(g)$$

(a) What is the mole ratio of O_2 to C_6H_6?
(b) How many moles of O_2 are required to react with each mole of C_6H_6?
(c) How many moles of O_2 are required to react with 0.38 mol of C_6H_6?

6.12 Aluminum reacts with chlorine gas according to the following equation:

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

(a) What is the mole ratio of Cl_2 to Al?
(b) How many moles of Cl_2 are required to react with each mole of Al?
(c) How many moles of Cl_2 are required to react with 0.085 mol of Al?

6.13 Suppose you have a piece of aluminum that you want to react completely by the following single-displacement reaction:

$$2Al(s) + 6HNO_3(aq) \longrightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$$

(a) What mole ratio would you use in the following equation to determine the number of moles of HNO_3 needed to react with a known amount of Al?

$$\text{mol Al} \times \frac{\qquad}{\qquad} = \text{mol } HNO_3$$

(b) If you add more than enough nitric acid so that all the aluminum reacts, what mole ratio would you use in the following equation to determine the moles of H_2 produced?

$$\text{mol Al} \times \frac{\qquad}{\qquad} = \text{mol } H_2$$

(c) Suppose you know the number of moles of H_2 product formed and you want to know the number of moles of Al that reacted. What mole ratio would you use in the following equation?

$$\text{mol } H_2 \times \frac{\qquad}{\qquad} = \text{mol Al}$$

6.14 Suppose you want to convert iron ore to a specific amount of pure iron using the following reaction:

$$Fe_3O_4(s) + 4CO(g) \longrightarrow 3Fe(s) + 4CO_2(g)$$

(a) What mole ratio would you use in the following equation to determine the number of moles of CO needed to react with a known amount of Fe_3O_4?

$$\text{mol } Fe_3O_4 \times \frac{\qquad}{\qquad} = \text{mol CO}$$

(b) If you add more than enough CO so that all the Fe_3O_4 reacts, what mole ratio would you use in the following equation to determine the moles of CO_2 produced?

$$\text{mol } Fe_3O_4 \times \frac{\qquad}{\qquad} = \text{mol } CO_2$$

(c) Suppose you know the number of moles of Fe product formed and you want to know the number of moles of CO that reacted. What mole ratio would you use in the following equation?

$$\text{mol Fe} \times \frac{\qquad}{\qquad} = \text{mol CO}$$

6.15 Consider the combustion reaction of propane, used in barbecue grills:

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

Which of the following is conserved in this reaction?
(a) moles of molecules
(b) moles of atoms
(c) atoms
(d) mass
(e) Which of your answers to (a) through (d) are true for any reaction?

6.16 Consider the combination reaction of hydrogen and iodine:

$$H_2(g) + I_2(s) \longrightarrow 2HI(g)$$

Which of the following is conserved in this reaction?
(a) moles of molecules
(b) moles of atoms
(c) atoms
(d) mass
(e) Which of your answers to (a) through (d) are true for any reaction?

6.17 The balanced equation for the decomposition of TNT, $C_7H_5(NO_2)_3$, is

$$2C_7H_5(NO_2)_3(s) \longrightarrow 7C(s) + 7CO(g) + 3N_2(g) + 5H_2O(g)$$

(a) Write the mole ratios that relate moles of each product to the reactant $C_7H_5(NO_2)_3$?
(b) How many moles of each product would form if 1.00 mol of $C_7H_5(NO_2)_3$ reacts?
(c) How many moles of each product would form if 6.25 mol of $C_7H_5(NO_2)_3$ react?

6.18 The balanced equation for the decomposition of ammonium nitrate is

$$2NH_4NO_3(s) \longrightarrow 2N_2(g) + O_2(g) + 4H_2O(g)$$

(a) Write the mole ratios that relate moles of each product to the reactant NH_4NO_3.
(b) How many moles of each product would form if 1.00 mol of NH_4NO_3 reacts?
(c) How many moles of each product would form if 12.4 mol of NH_4NO_3 react?

Mass-Mass Conversions

6.19 When copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$, is heated, it decomposes to the dehydrated form. The waters of hydration are released from the solid crystal and form water vapor. The hydrated form is medium blue, and the dehydrated solid is light blue. The balanced equation is

$$CuSO_4 \cdot 5H_2O(s) \xrightarrow{\text{heat}} CuSO_4(s) + 5H_2O(g)$$



(a) What is the molar mass of $CuSO_4 \cdot 5H_2O$?
(b) What is the molar mass of $CuSO_4$?
(c) If 1.00 g $CuSO_4 \cdot 5H_2O$ is decomposed to $CuSO_4$, predict the mass of the remaining light blue solid.

6.20 When the reddish-brown mercury(II) oxide, HgO, is heated, it decomposes to its elements, liquid mercury metal and oxygen gas:

$$2HgO(s) \xrightarrow{\text{heat}} 2Hg(l) + O_2(g)$$



(a) What is the molar mass of HgO?
(b) What is the molar mass of Hg?
(c) If 2.00 g HgO is decomposed to Hg, predict the mass of the pure Hg metal produced.

6.21 When active metals such as sodium are exposed to air, they quickly form a coating of metal oxide. The balanced equation for the reaction of sodium metal with oxygen gas is

$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

Suppose a piece of sodium metal gains 2.05 g of mass after being exposed to air. Assume that this gain can be attributed to its reaction with oxygen.
(a) What mass of O_2 reacted with the Na?
(b) What mass of Na reacted?
(c) What mass of Na_2O formed?

6.22 When aluminum metal is exposed to oxygen gas, a coating of aluminum oxide forms on the surface of the aluminum. The balanced equation for the reaction of aluminum metal with oxygen gas is

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

Suppose a sheet of pure aluminum gains 0.0900 g of mass when exposed to air. Assume that this gain can be attributed to its reaction with oxygen.
(a) What mass of O_2 reacted with the Al?
(b) What mass of Al reacted?
(c) What mass of Al_2O_3 formed?

6.23 Diiodine pentoxide is used in respirators to remove carbon monoxide from air:

$$I_2O_5(s) + 5CO(g) \longrightarrow I_2(s) + 5CO_2(g)$$

(a) What mass of carbon monoxide could be removed from air by a respirator that contains 50.0 g of diiodine pentoxide?
(b) What mass of I_2 would remain in the respirator?

6.24 Ethanol, used in alcoholic beverages, can be produced by the fermentation of sucrose, which is found in sugar cane and other plants. The balanced equation for the fermentation process is

$$C_{12}H_{22}O_{11}(s) + H_2O(l) \longrightarrow 4C_2H_5OH(l) + 4CO_2(g)$$

(a) What mass of ethanol, C_2H_5OH, would be produced when 2.50 g sucrose reacts by this process?
(b) What mass of CO_2 would also be produced?

6.25 For each of the following *unbalanced* equations, balance the equation and then determine how many grams of the second reactant would be required to react completely with 0.600 g of the first reactant.
(a) $Cr(s) + Cl_2(g) \longrightarrow CrCl_3(s)$
(b) $RbO_2(s) + H_2O(l) \longrightarrow O_2(g) + RbOH(s)$
(c) $C_5H_{12}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
(d) $Li(s) + Cl_2(g) \longrightarrow LiCl(s)$

6.26 For each of the following *unbalanced* equations, balance the equation and then determine how many grams of the first reactant would be required to react completely to give 1.70 g of the product indicated in boldface.
(a) $C_6H_{12}(l) + O_2(g) \longrightarrow \mathbf{CO_2}(g) + H_2O(g)$
(b) $NI_3(s) \longrightarrow N_2(g) + \mathbf{I_2}(s)$
(c) $Na(s) + O_2(g) \longrightarrow \mathbf{Na_2O}(s)$
(d) $Ca(s) + HCl(aq) \longrightarrow CaCl_2(aq) + \mathbf{H_2}(g)$

6.27 Calcium carbonate, $CaCO_3$, is often used in commercial antacids. It acts to reduce the acidity in the stomach by neutralizing stomach acid, which is mostly HCl, by the following reaction:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What mass of $CaCO_3$ is needed to neutralize 0.020 mol HCl?

6.28 Sulfuric acid is commonly used as an electrolyte in car batteries. Suppose you spill some on your garage floor. Before cleaning it up, you wisely decide to neutralize it with sodium bicarbonate (baking soda) from your kitchen. The reaction of sodium bicarbonate and sulfuric acid is

$$2NaHCO_3(s) + H_2SO_4(aq) \longrightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$$

You estimate that your acid spill contains about 2.0 mol H_2SO_4. What mass of $NaHCO_3$ do you need to neutralize the acid?

Limiting Reactants

6.29 When a fuel is burned in the air, what is typically the limiting reactant?

6.30 An acidic solution is added to a basic solution. If the resulting solution is basic, is the acid or the base the limiting reactant?

6.31 Suppose you are making turkey sandwiches for your friends. You have lots of friends, so you want to make as many sandwiches as possible. Because you are on a budget, you only have 24 pieces of bread and 15 pieces of turkey. Each sandwich consists of 2 pieces of bread and 1 piece of turkey.
(a) How many sandwiches can you make?
(b) What is the limiting part?
(c) What part is left over? How many are left over?

6.32 Your little sister is having a party and you are in charge of the party goody bags. You decide that each bag should have 3 candy bars, 1 bottle of nail polish, and 2 pieces of bubble gum. You bought an 18-pack of candy bars, 12 bottles of nail polish, and 20 pieces of bubble gum.
(a) What is the ratio of candy bars to bottles of nail polish to pieces of bubble gum in the goody bags?
(b) What is the limiting "goody"?
(c) Assuming you have an unlimited supply of bags, how many goody bags can you make?
(d) How many of each type of goody will be left over?

6.33 The balanced equation for the reaction of nitrogen and fluorine gas to form nitrogen trifluoride is

$$N_2(g) + 3F_2(g) \longrightarrow 2NF_3(g)$$

What is the limiting reactant when each of the following sets of quantities of reactants is mixed?
(a) 9 N_2 molecules and 9 F_2 molecules
(b) 5 N_2 molecules and 20 F_2 molecules
(c) 6 N_2 molecules and 18 F_2 molecules

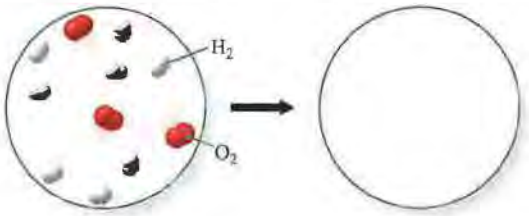
6.34 The balanced equation for the reaction of chromium metal and chlorine gas is

$$2Cr(s) + 3Cl_2(g) \longrightarrow 2CrCl_3(s)$$

What is the limiting reactant when each of the following sets of quantities of reactants is mixed?
(a) 4 Cr atoms and 6 Cl_2 molecules
(b) 6 Cr atoms and 10 Cl_2 molecules
(c) 12 Cr atoms and 20 Cl_2 molecules

6.35 The molecular-level diagram shows a mixture of reactant molecules (three O_2 molecules and eight H_2 molecules) for the following reaction:

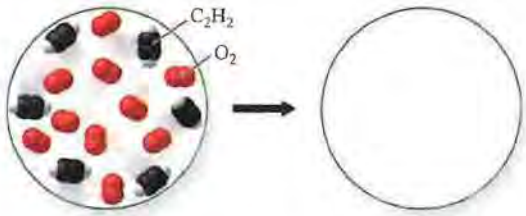
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$



(a) Draw what the mixture should look like when the reaction is complete.
(b) What is the limiting reactant?
(c) Which reactant is left over?

6.36 The molecular-level diagram shows a mixture of reactant molecules for the following reaction:

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$$



(a) Draw what the mixture should look like when the reaction is complete.
(b) What is the limiting reactant?
(c) Which reactant is left over?

6.37 Use the balanced equation for the combustion of ethane to complete the table.

	$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$			
Initially mixed	6 molecules	18 molecules	0 molecules	0 molecules
How much reacts			—	—
Composition of final mixture				

6.38 Use the balanced equation for the combustion of ethane to complete the table.

	$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$			
Initially mixed	10 molecules	20 molecules	0 molecules	0 molecules
How much reacts			—	—
Composition of final mixture				

6.39 Use the balanced equation for the combustion of butane to complete the table.

	$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$			
Initially mixed	3.10 mol	13.0 mol	0.00 mol	0.00 mol
How much reacts			—	—
Composition of final mixture				

6.40 Use the balanced equation for the combustion of butane to complete the table.

	$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$			
Initially mixed	0.500 mol	3.50 mol	0.00 mol	0.00 mol
How much reacts			—	—
Composition of final mixture				

6.41 The balanced equation for the reaction of phosphorus and oxygen gas to form tetraphosphorus decoxide is

$$P_4(g) + 5O_2(g) \longrightarrow P_4O_{10}(s)$$

What is the limiting reactant when each of the following sets of quantities of reactants is mixed?
(a) 0.50 mol P_4 and 5.0 mol O_2
(b) 0.20 mol P_4 and 1.0 mol O_2
(c) 0.25 mol P_4 and 0.75 mol O_2

6.42 Tetraphosphorus trisulfide, P_4S_3, is used in the heads of wooden matches. This material is manufactured by the heating of a mixture of red phosphorus and sulfur.

$$8P_4(s) + 3S_8(s) \longrightarrow 8P_4S_3(s)$$

What is the limiting reactant when each of the following sets of quantities of reactants is mixed?
(a) 4.4 mol P_4 and 1.1 mol S_8
(b) 0.20 mol P_4 and 0.10 mol S_8
(c) 0.75 mol P_4 and 0.28 mol S_8

6.43 The balanced equation for the reaction of nitrogen and fluorine to form nitrogen trifluoride is

$$N_2(g) + 3F_2(g) \longrightarrow 2NF_3(g)$$

Assume that 10.0 g of N_2 is mixed with 10.0 g of F_2.
(a) What is the limiting reactant?
(b) What is the maximum amount of NF_3, in grams, that can be produced?

6.44 The balanced equation for the reaction of aluminum metal and chlorine gas is

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

Assume that 0.40 g Al is mixed with 0.60 g Cl_2.
(a) What is the limiting reactant?
(b) What is the maximum amount of $AlCl_3$, in grams, that can be produced?

6.45 Sodium bicarbonate reacts with hydrochloric acid in a gas-forming reaction to produce aqueous sodium chloride, water, and carbon dioxide gas:

$$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

Determine the limiting reactant and mass of CO_2 gas produced when 8.0 g of $NaHCO_3$ is added to a solution that contains 4.0 g of HCl.

6.46 Sodium metal reacts with water in the following single-displacement reaction:

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Determine the limiting reactant and mass of H_2 gas produced when 2.0 g of Na is added to 10.0 g H_2O.

6.47 Use the balanced equation for the combustion of ethane to complete the table.

	$2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(g)$			
Initially mixed	0.260 g	1.00 g	0.00 g	0.00 g
How much reacts			—	—
Composition of final mixture				

6.48 Use the balanced equation for the combustion of butane to complete the table.

	$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$			
Initially mixed	0.130 g	0.580 g	0.00 g	0.00 g
How much reacts			—	—
Composition of final mixture				

6.49 An aqueous solution containing 10.0 g NaOH is added to an aqueous solution containing 10.0 g HNO_3.
(a) Write a balanced equation for the acid-base reaction that should occur.
(b) Identify the limiting reactant.
(c) Will the solution be acidic or basic when the reaction is complete?

6.50 An aqueous solution containing 10.0 g NaOH is added to an aqueous solution containing 10.0 g H_3PO_4.
(a) Write a balanced equation for the acid-base reaction that should occur.
(b) Identify the limiting reactant.
(c) Will the solution be acidic or basic when the reaction is complete?

Percent Yield

6.51 In a process for obtaining pure copper from copper ore, 55.6 g of copper metal was obtained. Is this the theoretical yield or actual yield?

6.52 After determining the limiting reactant in a reaction, you calculate the amount of product that can form from the limiting reactant. Is this the theoretical yield or the actual yield?

6.53 A student carried out the following precipitation reaction:

$$BaCl_2(aq) + Na_2CO_3(aq) \longrightarrow BaCO_3(s) + 2NaCl(aq)$$

After no more precipitate appeared to form, he filtered the barium carbonate solid and then immediately weighed it. Using this mass, he calculated the percent yield as 105%. Is this possible? Explain.

6.54 What factors may contribute to a percent yield less than 100%?

6.55 A student was synthesizing aspirin in the laboratory. Using the amount of limiting reactant, she calculated the mass of aspirin that should form as 8.95 g. When she weighed her aspirin product on the balance, its mass was 7.44 g.
(a) What is the actual yield of aspirin?
(b) What is the theoretical yield of aspirin?
(c) Calculate the percent yield for this synthesis.

6.56 A student added zinc metal to a copper nitrate solution to obtain copper metal by a single-displacement reaction. From the known moles of copper nitrate in solution, she calculated the expected mass of copper metal as 6.7 g. When she weighed the copper on a balance, the mass was 5.7 g.
(a) What is the actual yield of copper metal?
(b) What is the theoretical yield of copper metal?
(c) Calculate the percent yield.

6.57 The combination reaction of sodium metal and chlorine gas to form sodium chloride is represented by the following balanced equation:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

If 5.00 g of sodium metal reacts completely with excess chlorine gas, and 11.5 g NaCl is actually obtained, what is the percent yield of sodium chloride?

6.58 Ammonia is synthesized commercially from nitrogen gas and hydrogen gas for the production of fertilizers:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

If 100.0 g of nitrogen reacts completely with excess hydrogen, and 34.0 g of NH_3 are obtained, what is the percent yield of ammonia?

6.59 When I_2 is mixed with excess H_2, 0.80 mol HI is isolated in the laboratory. If this reaction occurs to give an 85% yield, how many moles of I_2 reacted? (Hint: Write a balanced equation.)

6.60 The reaction of lithium metal and water to form lithium hydroxide and hydrogen gas is represented by the following balanced equation:

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

When Li is mixed with excess water, 0.30 mol H_2 gas is isolated in the laboratory. If this reaction occurs to give an 85% yield of H_2, how many moles of Li reacted?

6.61 The combination reaction of magnesium metal and bromine to form magnesium bromide is represented by the following balanced equation:

$$Mg(s) + Br_2(l) \longrightarrow MgBr_2(l)$$

If 1.0 mol of Mg is mixed with 2.0 mol of Br_2, and 0.84 mol of $MgBr_2$ is obtained, what is the percent yield for the reaction?

6.62 The combination reaction of sodium metal and nitrogen gas to form sodium nitride is represented by the following balanced equation:

$$6Na(s) + N_2(g) \longrightarrow 2Na_3N(s)$$

If 0.30 mol Na is mixed with 0.60 mol N_2, and 0.092 mol Na_3N is obtained, what is the percent yield for the reaction?

Energy Changes

6.63 If energy cannot be created or destroyed, what happens to the energy of a ball as it rolls down a hill and rests at the bottom?

6.64 If energy cannot be created, what is the source of the energy that is released when gasoline is burned?

6.65 Consider the reaction of hydrogen and oxygen in a hydrogen-fuel vehicle. Describe the transformation of energy in this process in terms of potential and kinetic energy.

6.66 Consider propane fuel being burned in a propane-fueled barbecue grill. Describe the transformation of energy in this process in terms of potential and kinetic energy.

6.67 When the solids ammonium thiocyanate and barium hydroxide are mixed in a beaker, a solution forms, and the temperature of the solution drops to –5°C. Is this an endothermic or exothermic reaction? Explain.

6.68 When baking soda is added to a solution of sulfuric acid, the container becomes warmer to the touch. Is this an endothermic or exothermic reaction? Explain.

6.69 A doctor sprays a liquid on your skin to numb it before injecting a painkiller. You notice that the liquid evaporates quickly. You also notice that your skin feels cold. Explain why your skin feels cold.

6.70 Explain why sweating is a natural way for your body to cool itself.

6.71 What is the relationship between the potential energy of the reactants and products in an exothermic reaction?

6.72 If the reactants have a lower potential energy than the products, is the reaction exothermic or endothermic?

6.73 Convert an energy of 526 cal to units of joules.

6.74 Convert an energy of 225 cal to units of joules.

6.75 Convert an energy of 145 kJ to units of calories.

6.76 Convert an energy of 0.675 kJ to units of calories.

6.77 Convert an energy of 876 J to units of Calories.

6.78 Convert an energy of 2430 J to units of Calories.

6.79 Sodium cyclamate, $NaC_6H_{12}NSO_3$, was a popular nonsugar sweetener until it was banned by the Food and Drug Administration. Its sweetness is about 30 times that of sugar (or sucrose, $C_{12}H_{22}O_{11}$). The energy that can be supplied by ingestion of each compound is 16.03 kJ/g for sodium cyclamate and 16.49 kJ/g for sucrose. What is the energy savings in Calories resulting from using 1.00 g of sodium cyclamate in place of 30.0 g of sucrose?

6.80 In some areas, gasoline is formulated to contain 15% ethanol during some seasons. If gasoline releases 11.4 kcal/g upon combustion, but ethanol releases only 7.12 kcal/g, how much less energy in kJ/g is provided by the mixture than by pure gasoline?

6.81 If the same amount of heat is added to 50.0 g each of the metals, aluminum, copper, and lead, all at the same initial temperature, which metal will have the highest final temperature? Refer to Table 6.2 for specific heat values.

6.82 We add the same amount of heat to three different samples of the same liquid substance. The temperature of the first sample increases in temperature by the greatest amount. How can we explain the different temperature changes in each sample?

6.83 How much heat has to be added to 528 g of copper at 22.3°C to raise the temperature of the copper to 49.8°C? (See Table 6.2.)

6.84 How much heat has to be added to 235 g of iron at 25.0°C to raise the temperature of the iron to 250.0°C? (See Table 6.2.)

6.85 What is the heat change when 1.25 g of water vapor (steam) at 185.3°C is cooled to 102.1°C? The specific heat of steam is 2.02 J/(g °C).

6.86 What is the heat change when 55.0 g of water cools from 60.0°C to 25.5°C?

6.87 What are some characteristics of a good calorimeter?

6.88 What information can be obtained from a calorimetry experiment?

6.89 A rock is added to 60.0 g of water. The initial temperature of the water is 25.0°C. The final temperature of the water and the rock is 30.1°C.
(a) Did the rock gain or lose heat?
(b) What is the heat change of the rock?

6.90 A marble is added to 40.0 g of water. The initial temperature of the water is 25.0°C. The final temperature of the water and the marble is 22.2°C.
(a) Did the marble gain or lose heat?
(b) What is the heat change of the marble?

6.91 A piece of copper is placed in 75.0 g of water held in a calorimeter at 45.0°C. The final temperature of the water is 36.2°C.
(a) Is the initial temperature of the copper greater than or less than the initial temperature of the water?
(b) Did the copper gain or lose heat?
(c) What is the heat change of the copper?

6.92 A scalpel is heated in an autoclave and then dropped into 145.0 g of water held in a calorimeter at 22.0°C. The final temperature of the water is 33.1°C.
(a) Did the scalpel gain or lose heat?
(b) Did the water gain or lose heat?
(c) What is the heat change of the scalpel?

6.93 What is the final temperature when 70.0 g of water at 80.0°C is mixed with 30.0 g of water at 25.0°C?

6.94 What is the final temperature when a 100.0-g sample of copper pellets at 100.0°C is added to a calorimeter containing 140.0 g of water at 25.0°C? Refer to Table 6.2 for specific heat values.

Heat Changes in Chemical Reactions

6.95 Why can't you measure the temperature change of a chemical reaction?

6.96 Why do you have to first determine the heat change of the surroundings for a chemical reaction?

6.97 You carry out a chemical reaction in a calorimeter to determine the heat change for the reaction. Identify the system and surroundings.

6.98 When a chemical reaction occurs in a calorimeter containing water, and the temperature of the water decreases, is the reaction endothermic or exothermic?

6.99 A mixture of carbon monoxide, hydrogen, and oxygen gases, known as *water gas*, can be used as an industrial fuel. The reaction of these gases is

$$CO(g) + H_2(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

This reaction releases 525 kJ of heat for every mole of carbon monoxide that reacts.
(a) Is this reaction endothermic or exothermic?
(b) What is the energy change for this reaction in units of kJ/mol of carbon monoxide?

6.100 The reaction that produces the *water gas* mixture, described in Question 6.99, is

$$C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)$$

This reaction requires an input of 131 kJ of heat for every mole of carbon that reacts.
(a) Is this reaction endothermic or exothermic?
(b) What is the energy change for this reaction in units of kJ/mol of carbon?

6.101 When a 6.00-g sample of coal is burned, it releases enough heat to raise the temperature of 2010 g of water from 24.0°C to 41.5°C.
(a) How much heat did the coal release as it burned?
(b) Calculate the heat of combustion of coal in units of kJ/g.

6.102 When a 4.00-g sample of magnesium metal is burned, it produces enough heat to raise the temperature of 2010 g of water from 24.00°C to 33.10°C.
(a) How much heat did the magnesium release as it burned?
(b) Calculate the heat of combustion of magnesium in units of kJ/g.

6.103 A 2.00-g peanut is burned in a bomb calorimeter containing 1200 g of water. The temperature of the water increases from 25.00°C to 30.25°C.
(a) How much heat, in joules, did the peanut release as it burned?
(b) Calculate the heat content in units of calories and Calories.
(c) Calculate the energy value in units of Cal/g.

6.104 A 5.00-g piece of fat is burned in a bomb calorimeter containing 4050 g of water. The temperature of the water increases by 12.4°C.
(a) How much heat, in joules, did the fat release as it burned?
(b) Calculate the heat content in units of calories and Calories.
(c) Calculate the energy value in units of Cal/g.
(d) Calculate the heat content in units of Cal/mol. Assume the fat is all tristearin, a typical fat with the molecular formula $C_{57}H_{110}O_6$.

6.105 The heat change that accompanies the formation of 1.00 mol of carbon dioxide from its elements is −393.7 kJ/mol. What heat change accompanies the formation of 0.650 mol CO_2?

6.106 The decomposition of hydrogen peroxide to form water and oxygen gas releases 196.6 kJ per mole of hydrogen peroxide. This reaction occurs when hydrogen peroxide is placed on a cut to sterilize it. How much heat is released when 5.25 mol H_2O_2 decompose?

Additional Questions

6.107 When nitroglycerin explodes, it decomposes to form carbon dioxide gas, nitrogen gas, oxygen gas, and water vapor. The balanced equation is

$$4C_3H_5O_9N_3(l) \longrightarrow 12CO_2(g) + 6N_2(g) + O_2(g) + 10H_2O(g)$$

(a) If 1.00 mol $C_3H_5O_9N_3$ decomposes, how many moles of each gaseous product should form?
(b) If 2.50 mol $C_3H_5O_9N_3$ decomposes, how many moles of each gaseous product should form?

6.108 Certain drain cleaners are a mixture of sodium hydroxide and powdered aluminum. When dissolved in water, the sodium hydroxide reacts with the aluminum and the water to produce hydrogen gas.

$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \longrightarrow 2NaAl(OH)_4(aq) + 3H_2(g)$$

The sodium hydroxide helps dissolve grease, and the hydrogen gas provides a mixing and scrubbing action. What mass of hydrogen gas would be formed from a reaction of 2.48 g Al and 4.76 g NaOH in water?

6.109 When silver nitrate is added to an aqueous solution of calcium chloride, a precipitation reaction occurs that removes the chloride ions from solution.

$$2AgNO_3(s) + CaCl_2(aq) \longrightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

(a) If a solution contains 10.0 g $CaCl_2$, what mass of $AgNO_3$ should be added to remove all of the chloride ions from solution?

(b) When enough $AgNO_3$ is added so that all 10.0 g of $CaCl_2$ react, what mass of the AgCl precipitate should form?

6.110 Sometimes it is more convenient to measure the amount of a liquid reactant in volume units instead of mass units. Given that 1.00 gal is 3.79 L and the density of methanol is 0.793 g/mL, what mass of oxygen is needed to react with 1.00 gal of methanol? The balanced equation is

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(g)$$

6.111 The balanced equation for the combustion of octane is

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

What mass of oxygen is needed to react with 1.00 gal of octane? (A gallon is 3.79 L, and the density of octane is 0.703 g/mL.)

6.112 When lead(II) nitrate is added to an aqueous solution of sodium iodide, a precipitation reaction occurs that removes the iodide ion from solution.

$$Pb(NO_3)_2(s) + 2NaI(aq) \longrightarrow 2PbI_2(s) + 2NaNO_3(aq)$$

(a) If a solution contains 7.00 g NaI, what mass of $Pb(NO_3)_2$ should be added to remove all of the iodide ion from solution?

(b) When enough $Pb(NO_3)_2$ is added so that all 7.00 g of NaI reacts, what mass of PbI_2 should precipitate?

6.113 A 5.0-g sample of greenish-yellow Cl_2 gas is added to a 10.0-g sample of gray potassium metal to form a white solid in a sealed container.

(a) Write a balanced equation for the combination reaction that should occur.

(b) Identify the limiting reactant.

(c) Predict the appearance of the substances left in the reaction container once the reaction is complete.

(d) Calculate the mass of product that should form.

(e) What mass of excess reactant should be mixed with the product after the reaction is complete?

6.114 A 5.00-g piece of copper metal is added to a colorless aqueous solution that contains 20.0 g of $Zn(NO_3)_2$. The products that form are a water-soluble blue ionic compound and a gray metal.

(a) Write a balanced equation for the single-displacement reaction that should occur.

(b) Identify the limiting reactant.

(c) Predict the appearance of the substances left in the reaction container once the reaction is complete.

6.115 The equation for the combustion of hydrogen is

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

(a) What is the mole ratio of O_2 to H_2?

(b) If 1.0 g of H_2 reacts, what mass of O_2 will react with it, and what mass of H_2O should form?

(c) When 1.0 g H_2 is mixed with 4.0 g O_2, what is the theoretical yield of H_2O?

6.116 The decomposition reaction of calcium carbonate is represented by the following balanced equation:

$$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

After a 15.8-g sample of calcium carbonate was heated in an open container to cause decomposition, the mass of the remaining solid was determined to be 9.10 g. The student is unsure if the reaction is complete, so the solid could contain unreacted $CaCO_3$.

(a) Determine the theoretical yield of CaO and the theoretical yield of CO_2.

(b) Determine the percent yield of CaO and the percent yield of CO_2.

(c) Why are the percent yields for CaO and CO_2 different?

6.117 What is the temperature of a 125-g sample of tin, initially at 26.2°C, if 155 J of heat is added to the tin? See Table 6.2.

6.118 Calculate the temperature change that results when 2.15 kJ of heat is supplied to 45.0 g of aluminum. See Table 6.2.

6.119 If 20.0 g of an unknown metal cools from 50.0°C to 25.5°C and loses 220.6 J of heat, what is the specific heat of the metal in joules per gram per degree Celsius? Refer to Table 6.2 to determine the possible identity of the metal.

6.120 What is the specific heat of an unknown metal if 658 J of heat increases the temperature of 1.00 kg of the metal by 5.10°C? Refer to Table 6.2 to determine the possible identity of the metal.

6.121 When 0.250 mol of glucose is broken down in your body, 701 kJ of energy is released. What is the energy change in units of kJ/mol and Cal/mol?

6.122 A cold object at 5°C is placed in an insulated cup of water at 25°C. Determine whether or not each of the following will occur.
(a) The increase in temperature of the object will be equal to the decrease in temperature of the water.
(b) Only the object will change temperature.
(c) Only the water will change temperature.
(d) The heat gain by the object will equal the heat lost by the water.
(e) The final temperature will be somewhere between 5°C and 25°C.

6.123 If 20.0 g of lead at 24.5°C is added to 105 g of water at 55.0°C, what will be the final temperature of the mixture? See Table 6.2.

6.124 Methylhydrazine, CH_6N_2, is commonly used as a liquid rocket fuel. The heat of combustion of methylhydrazine is -1.30×10^3 kJ/mol. How much heat is released when 100.0 g of methylhydrazine is burned?

6.125 A 175-g pipe is heated to 78.24°C. The pipe is then placed in 100.0 g of water held in a calorimeter at 25.00°C. The final temperature of the mixture is 33.43°C. The specific heat of water is 4.184 J/(g °C). Assuming no heat is transferred out of the water, what is the specific heat of the pipe? Using Table 6.2, determine what the pipe could be made of.

6.126 A 150.0-g sample of copper is heated to 89.3°C. The copper is then placed in 125.0 g of water held in a calorimeter at 22.5°C. The final temperature of the mixture is 29.0°C. Assuming no heat is lost from the water, what is the specific heat of copper?

CHAPTER

13 Acids and Bases

13.1 What Are Acids and Bases?

13.2 Strong and Weak Acids and Bases

13.3 Relative Strengths of Weak Acids

13.4 Acidic, Basic, and Neutral Solutions

13.5 The pH Scale

13.6 Buffered Solutions

Summary

Math Toolbox 13.1 Log and Inverse Log Functions

Key Relationships

Key Terms

Questions and Problems



Olivia, Jake, and Sara are college students who have part-time jobs during the academic year. Olivia works for a plant nursery where she is in charge of watering and fertilizing perennials. She likes her job because she is a natural at gardening and has taken care of plants since she was a child. She also likes the challenge of figuring out how to help plants thrive. Jake handles customer service and pool maintenance for a swimming pool company. He tests water for several regular customers and adds the right chemicals to maintain safe and attractive pools. Sara cycles to a nearby candy factory on weekends where she helps make and package sour candies. Sara enjoys the free samples, but she also likes dropping in at the plant's quality-control lab where product samples are tested.

What do Olivia's, Jake's, and Sara's jobs have in common? Each job uses chemistry, especially the chemistry of acids and bases. Olivia's job at the nursery requires her to measure the pH of plant soil solution, the water, and dissolved substances surrounding the plants she tends. Many minerals and organic compounds in soil are acids or bases that can influence pH, and soil pH affects which metal ions dissolve in the soil solution. If the soil is too acidic ($pH < 6$), Olivia must add lime, CaO, a base that reduces the acidity level of the soil. She's careful to add the correct amount of lime. Too much makes the soil too basic ($pH > 8$). Under basic conditions, some nutrients form insoluble compounds that plants cannot absorb.

Olivia also educates home gardeners about how to maintain healthy soil in their yards. Garden soil can become too acidic for many reasons. One cause is the use of acidifying nitrogen fertilizers. Many popular brands contain the ammonium ion, NH_4^+, which reacts as an acid with soil bacteria. Another cause of acidic soil is *acid rain,* which has a pH below 5.5. Acid rain is produced when sulfur and nitrogen oxides in the air react with water. In acidic soil more nutrients dissolve into the soil solution. Important nutrients may get washed away, or their concentrations may increase to levels toxic to plants.

When Jake maintains pools, he uses skimmers and filters to remove suspended particles such as leaves, dirt, and hair from the water. He also treats the water to remove—or render harmless—dissolved pollutants such as body wastes, algae, and disease-causing bacteria. He uses chlorine in a variety of forms to kill bacteria. When the concentration of effective chlorine in a pool falls too low, Jake superchlorinates (*shocks*) the pool water by adding calcium hypochlorite, $Ca(OCl)_2$, a source of chlorine. The compound dissolves in water to form $OCl^-(aq)$ ions, which react with water to form hypochlorous acid, $HOCl(aq)$, a strong disinfectant. The HOCl kills bacteria by penetrating their cell walls and destroying cellular proteins.

The pH level of the pool affects the relative amounts of $HOCl(aq)$ and $OCl^-(aq)$ in the water. At a pH of 7.5, the concentration of HOCl is most effective. If the pH falls below 7.2, the pool is too acidic. There is too much HOCl in the pool, and its high levels irritate the eyes of swimmers and can etch plaster walls. Above a pH of 7.8, there is not enough HOCl, and scale formation can occur, and the pool may become cloudy. Jake measures the pH of the pool regularly as part of his maintenance schedule (Figure 13.1). Usually he uses pH indicators that, when added to a sample of pool water, indicate pH by color. Sometimes he uses a portable pH meter that reads the pH directly. If the pH is too high, the pool water is too basic, so he must add an acid to lower the pH. Hydrochloric acid, HCl, also called muriatic acid, lowers the pH. If the pH is too low (too acidic), the base sodium carbonate, Na_2CO_3, is added to raise it.

Sara's job at the candy factory also involves acids (Figure 13.2). Sour candies contain a mixture of sugar and the organic acids malic acid and citric acid. Malic acid gives apples their "bite," and citric acid causes the sour taste of lemons and limes. All acids taste sour, although the degree of sourness varies from acid to acid. In addition to sour candy, many foods such as fruits contain acids. Vinegar, which is



FIGURE 13.1 Acid-base chemistry is important to a healthy swimming pool.

The gases NO, NO_2, SO_2, and SO_3 react with water in clouds to produce the acids HNO_3, H_2SO_3, and H_2SO_4. The result is acid rain.

The chlorine in swimming pool water is not elemental chlorine, Cl_2. If Cl_2 was added to pool water, it would not stay as Cl_2. It would react with water to form hypochlorous acid, HOCl, and hydrochloric acid, HCl. Because Cl_2 is a gas and very toxic, it is a very dangerous chemical to handle. Other products such as "liquid" chlorine (chlorine dissolved in water) or solid compounds such as NaOCl (bleach) or $Ca(OCl)_2$ are used instead.



FIGURE 13.2 The taste of sour candy can be attributed to the presence of acids.



FIGURE 13.3 Many household products contain acids or bases. Which of these products do you think contain acids and which contain bases?

Make a list of the acids and bases you encounter in your daily life.

found in pickles, salad dressings, and pizza sauce, contains acetic acid. Carbonated sodas contain phosphoric acid, carbonic acid, and sometimes citric acid. The next time you look at a food label, check to see if any acids are listed as ingredients.

While many foods contain acids, bases are rarely found in foods. One reason is that bases taste bitter. Many plants, some poisonous, contain organic bases called alkaloids. Their bitter taste is nature's way of telling us not to eat them. Some bases, however, are useful (Figure 13.3). Antacids, for example, are bases that act to neutralize excess acid in the stomach. While our stomachs can usually tolerate a high concentration of stomach acid, HCl, we sometimes secrete too much and get heartburn. Commonly used antacids contain magnesium hydroxide, $Mg(OH)_2$, aluminum hydroxide, $Al(OH)_3$, or calcium carbonate, $CaCO_3$. Bases are also found in some household cleaners. Ammonia, NH_3, is found in window cleaners, and sodium hydroxide, NaOH, is found in many oven and drain cleaners.

Some of the properties of acids and bases have been known for millennia, but only in the past 100 years have chemists come to understand the behavior of acids and bases at the molecular level. In this chapter we will explore the behavior of acids and bases in detail. We will look at the origin of the pH scale and find out how acid-base buffers work to maintain constant pH levels in solutions such as human blood.

Questions for Consideration

13.1 How do acids and bases differ from other substances?
13.2 How do strong and weak acids differ?
13.3 How can we compare the different strengths of weak acids?
13.4 What causes aqueous solutions to be acidic or basic?
13.5 How is pH related to acidity and basicity?
13.6 What is a buffered solution?

Math Tools Used in This Chapter

Scientific Notation (Math Toolbox 1.1)
Log and Inverse Log Functions (Math Toolbox 13.1)

13.1 | WHAT ARE ACIDS AND BASES?

Although there was a time when chemists routinely tasted any new compound that they made, we now know better. Never taste a chemical.

Centuries ago substances were first classified as acids and bases by their observable properties. Like the acids in sour candy or lemons, all acids have a sour taste. They are also corrosive to most metals, so people learned a long time ago not to store fruit juices or vinegar in metal containers. In contrast to acids, bases taste bitter and have a slippery feel. Acids and bases change the color of some dyes. For example, bases turn red litmus dye to blue. Acids change it back to red. When acids and bases react in neutralization reactions, their acidic and basic properties disappear.

Acid and Base Definitions

In the late 1800s a Swedish doctoral student named Svante Arrhenius proposed a definition of acids and bases based on his experiments with electrolytes. Because aqueous solutions of acids and bases conduct electricity, Arrhenius knew that the compounds were forming positive and negative ions in solution. Arrhenius proposed that acids and bases produce specific types of ions that other compounds do not. In the **Arrhenius model of acids and bases**, an acid in aqueous solution produces hydrogen ions, H^+, and a base produces hydroxide ions, OH^-. For example, hydrochloric acid ionizes in water to form $H^+(aq)$ and $Cl^-(aq)$:

$$HCl(g) \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

The base sodium hydroxide dissociates when it dissolves in water to form sodium ions and hydroxide ions:

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

The Arrhenius model explains how acids and bases neutralize each other. In a neutralization reaction of HCl and NaOH, for example, the H^+ from the HCl combines with the OH^- from NaOH to produce water:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Although his model does not adequately describe acids and bases as we know them today, Arrhenius earned the 1903 Nobel Prize in Chemistry because he insisted that the $H^+(aq)$ and $OH^-(aq)$ ions are important in acid-base behavior. Because of the simplicity of its definition, the model is still used when only a brief explanation of acid-base behavior is needed.

The Arrhenius definition of acids and bases is limited by some fundamental problems. The hydrogen ion, H^+, essentially a proton with an extremely small radius, has a very concentrated positive charge. As a result, free H^+ ions are unlikely to exist in aqueous solution. Instead they are associated with surrounding water molecules. Chemists today normally represent the aqueous H^+ ion as $H_3O^+(aq)$, the **hydronium ion** (Figure 13.4). Another limitation to the Arrhenius model is that it assumes all bases contain OH^- ions, but some do not. Many other ionic compounds, commonly called salts, have basic properties, such as the ability to neutralize acids. Examples of salts that are bases are metal oxides, carbonates, and fluorides. Some molecular compounds that contain no –OH group, such as ammonia, NH_3, also have basic properties.

In 1923 J. N. Brønsted, a Danish chemist, and T. M. Lowry, a British chemist, independently proposed a new theory to deal with the limitations of the Arrhenius model. The **Brønsted-Lowry theory** defines an acid as any substance that can donate an H^+ ion to another substance; it defines a base as any substance that can accept an H^+ ion from another substance. This definition includes all the Arrhenius acids and bases plus some ions in salts and some molecular compounds such as ammonia. According to the Brønsted-Lowry theory, when HCl gas dissolves in water, it donates an H^+ to H_2O to form hydronium ions and chloride ions:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Note that H_2O acts as a Brønsted-Lowry base by accepting the H^+ from HCl.

Ammonia dissolved in water acts as a Brønsted-Lowry base by the following equation:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

In this case, water acts as an acid by donating the H^+ to the NH_3 molecule. Note that $OH^-(aq)$ ions form, but the –OH group is not present as part of the base initially. Hydroxide ions form as a result of the Brønsted-Lowry acid-base reaction between NH_3 and H_2O. There are many organic amines (compounds that contain the $-NH_2$ group) that act as bases in water. An example is methylamine, CH_3NH_2:

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$$

Brønsted-Lowry bases also include anions, such as the carbonate ion, CO_3^{2-}, which is found in the compound Na_2CO_3 used to alter the pH of swimming pool water. When added to water, Na_2CO_3 first dissociates completely into its ions:

$$Na_2CO_3(s) \xrightarrow{H_2O} 2Na^+(aq) + CO_3^{2-}(aq)$$

ANIMATION: An Acid-Base Reaction

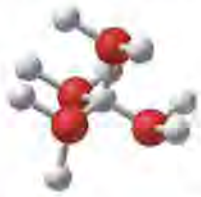


FIGURE 13.4 The aqueous H^+ ion is strongly associated with water molecules, most commonly surrounded by four molecules, as shown here, giving a formula of $H_9O_4^+(aq)$. Chemists, however, generally represent aqueous H^+ as the hydronium ion, $H_3O^+(aq)$.

In future courses you might encounter another definition of an acid or base that is named for G. N. Lewis. He defined an acid as an electron pair acceptor and a base as an electron pair donor. This broader definition encompasses many substances that are not explained by the Brønsted-Lowry theory but is beyond the scope of this text.

Recall from Chapter 12 that equilibrium arrows ($\rightleftharpoons$) show a system at equilibrium. The forward and reverse reactions occur at equal rates when the system is at equilibrium.

Some Brønsted-Lowry acid-base reactions take place without water. For example, ammonia gas reacts with hydrogen chloride gas to form ammonium ions and chloride ions. The ammonium and chloride ions immediately combine to form the solid ionic compound, ammonium chloride, $NH_4Cl(s)$.

$$\underset{\text{base}}{NH_3(g)} + \underset{\text{acid}}{HCl(g)} \longrightarrow NH_4Cl(s)$$



The CO_3^{2-} ion acts as a base in water by accepting an H^+ from H_2O:

$$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$$

Brønsted-Lowry acid-base reactions do not always involve water as the acid or base. Sometimes another substance in the water is better able to give or receive an H^+ ion. In the following reaction, hydrofluoric acid donates an H^+ to the carbonate ion base to form fluoride ion and bicarbonate ion:

$$\underset{\text{acid}}{HF(aq)} + \underset{\text{base}}{CO_3^{2-}(aq)} \rightleftharpoons F^-(aq) + HCO_3^-(aq)$$

Use Example 13.1 to practice identifying the Brønsted-Lowry acid and base in acid-base reactions.

EXAMPLE 13.1 | Brønsted-Lowry Acids and Bases

For each reaction, identify the Brønsted-Lowry acid and base reactants.

(a) $HNO_3(l) + H_2O(l) \longrightarrow NO_3^-(aq) + H_3O^+(aq)$
(b) $SO_4^{2-}(aq) + H_2O(l) \rightleftharpoons HSO_4^-(aq) + OH^-(aq)$
(c) $H_2CO_3(aq) + CH_3NH_2(aq) \rightleftharpoons HCO_3^-(aq) + CH_3NH_3^+(aq)$

Solution:

(a) The name of HNO_3, nitric acid, tells us that it is an acid. However, we can also identify it as an acid because it donates an H^+ to water as shown in the equation. Water accepts H^+ from HNO_3 to form $H_3O^+(aq)$, so H_2O is the base.
(b) The sulfate ion, SO_4^{2-}, does not have any hydrogens to donate, so it cannot be the acid. Instead, it is the base because it accepts H^+ to form $HSO_4^-(aq)$. Water donates the H^+, forming $OH^-(aq)$, so H_2O is the acid in this reaction.
(c) Methylamine, CH_3NH_2, acts as a base in water because it accepts an H^+, forming $CH_3NH_3^+(aq)$. Carbonic acid, H_2CO_3, donates the H^+, forming $HCO_3^-(aq)$, so H_2CO_3 is the acid in this reaction.

Practice Problem 13.1

For each reaction, identify the Brønsted-Lowry acid and base reactants.

(a) $OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$
(b) $H_2SO_4(aq) + F^-(aq) \longrightarrow HSO_4^-(aq) + HF(aq)$
(c) $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

Further Practice: Questions 13.15 and 13.16 at the end of the chapter

Conjugate Acid-Base Pairs

When an acid donates an H^+ to a base, the two products differ from the reactants by one H^+ ion. The product that forms as a result of gaining an H^+ ion is called the **conjugate acid** of the base from which it forms. The product that forms as a result of losing an H^+ ion is called the **conjugate base** of the acid from which it forms. Consider the reaction of hydrogen chloride gas with water:

$$\underset{\text{acid}}{HCl(g)} + \underset{\text{base}}{H_2O(l)} \longrightarrow \underset{\text{conjugate acid}}{H_3O^+(aq)} + \underset{\text{conjugate base}}{Cl^-(aq)}$$

In water, the acid, HCl, reacts to form its conjugate base, $Cl^-(aq)$. The base, H_2O, reacts to form its conjugate acid, $H_3O^+(aq)$. The acid reactant and the conjugate base product are a *conjugate acid-base pair.* Likewise, the base reactant and the conjugate acid product are a conjugate acid-base pair. Conjugate acid-base pairs always differ by one H^+ ion. The conjugate acid of a substance has one more H^+; the conjugate base of a substance has one less H^+. The conjugate base of an acid is what is left after the acid donates an H^+ to another substance. The conjugate base, then, has one less H atom in its formula and a decrease in charge of 1. Given that a substance reacts as an acid or a base, you should be able to identify its conjugate.

The strengths of conjugate bases of acids (or conjugate acids of bases) vary considerably. For example, the conjugate bases of strong acids are so weak that they do not react with water at all. We'll explore the strengths of conjugate acids and bases later in this chapter.

EXAMPLE 13.2 | Conjugate Acids and Bases

Identify the conjugate base for each acid and explain its charge.

(a) HOCl
(b) $H_2PO_4^-$
(c) H_2O

Solution:

(a) The conjugate base of HOCl is OCl^-. It has one less H atom than the acid and a charge of 1–, one less than the zero charge on HOCl.
(b) The conjugate base of $H_2PO_4^-$ is HPO_4^{2-}, which has one less H atom than the acid and a charge of 2–, one less than the 1– charge on $H_2PO_4^-$.
(c) The conjugate base of H_2O is OH^-, which has one less H atom than H_2O and a charge of 1–, one less than the zero charge on H_2O.

Practice Problem 13.2

Identify the conjugate acid for each base.

(a) F^-
(b) HCO_3^-
(c) H_2O

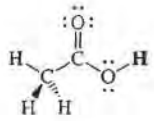
Further Practice: Questions 13.17 and 13.18 at the end of the chapter

In some of the examples we have looked at, water acts as an acid, while in others t acts as a base. Water can act as an acid by donating an H^+ to another substance, orming hydroxide ion, $OH^-(aq)$. It can also act as a base by accepting an H^+ from nother substance, forming hydronium ion, $H_3O^+(aq)$. The conjugate of water, ierefore, depends on whether water is acting as an acid or a base in a particular :action. A substance that can act as either an acid or a base is called an **ampho-:ric substance**. Water is the most common amphoteric substance. Another is the icarbonate ion, HCO_3^-. This anion is found in sodium bicarbonate, which is used) neutralize both acids and bases. When mixed with a basic solution containing $H^-(aq)$, the bicarbonate ion acts as an acid and its conjugate base is CO_3^{2-}:

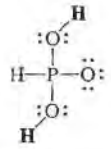
$$\underset{\text{acid}}{HCO_3^-(aq)} + OH^-(aq) \longrightarrow CO_3^{2-}(aq) + H_2O(l)$$

hen mixed with an acidic solution containing $H_3O^+(aq)$, the bicarbonate ion acts a base and its conjugate acid is H_2CO_3:

$$\underset{\text{base}}{HCO_3^-(aq)} + H_3O^+(aq) \longrightarrow H_2CO_3(aq) + H_2O(l)$$



acetic acid



phosphorous acid

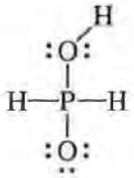
FIGURE 13.5 In oxoacids, the acidic hydrogens are bonded to oxygen atoms. How many acidic hydrogens do these acids have?

Acidic Hydrogen Atoms

In acids that contain more than one hydrogen atom, how do we know which hydrogen atoms are acidic? In other words, which hydrogen atoms ionize in water? An example is acetic acid, a carboxylic acid that is commonly represented with the formula CH_3CO_2H. Its formula shows that each molecule contains four hydrogen atoms. Which of the four are acidic? The Lewis structure of acetic acid is shown in Figure 13.5. From its Lewis structure we see that one hydrogen atom is bonded to an oxygen atom, and the other three are bonded to a carbon atom. Only the hydrogen atom bonded to an oxygen atom is acidic; it is the only one that ionizes when acetic acid dissolves in water. The three hydrogen atoms bonded to carbon are not acidic. When there is more than one hydrogen atom in an oxoacid, the acidic hydrogen atoms are usually bonded to oxygen. This is the case for phosphorous acid, H_3PO_3. Its Lewis structure (Figure 13.5) shows that two of the hydrogen atoms are bonded to oxygen atoms. They are the acidic hydrogen atoms. The hydrogen atom bonded to the phosphorus atom is not acidic. It is sometimes difficult to determine the number of acidic hydrogen atoms from the molecular formula of an acid, but it's generally easier when using a Lewis structure or molecular model.

EXAMPLE 13.3 | Identifying Acidic Hydrogens

Identify the acidic hydrogen atom(s) in hypophosphorous acid, H_3PO_2.

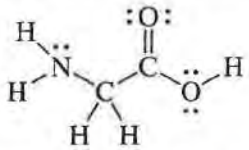


Solution:

Hypophosphorous acid, H_3PO_2, has only one acidic hydrogen. It is the hydrogen bonded to the oxygen atom.

Practice Problem 13.3

Identify the acidic hydrogen atom(s) in the amino acid glycine.

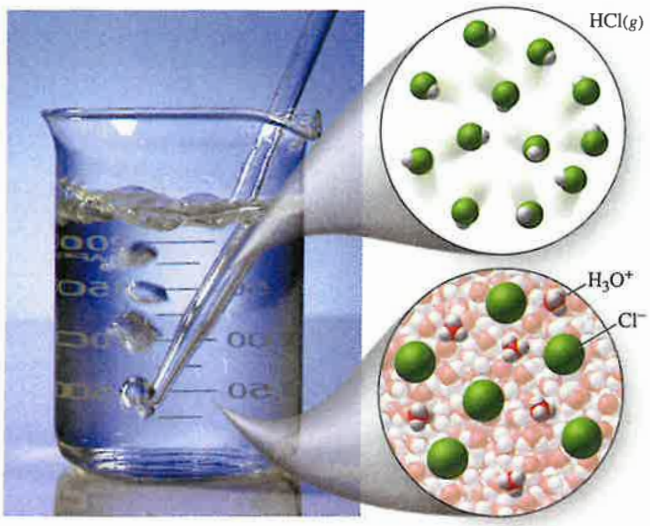


Further Practice: Questions 13.25 and 13.26 at the end of the chapter

13.2 STRONG AND WEAK ACIDS AND BASES

Jake uses a solution of HCl to lower the pH of pool water that is too basic. One reason that he uses HCl is that it is a strong electrolyte that ionizes completely in solution. As described in Chapter 11, some acids and bases are strong electrolytes, and others are weak electrolytes. An acid or base that is a strong electrolyte and *completely* ionizes or dissociates in water is a **strong acid** or a **strong base**. An acid or a base that is a weak electrolyte and only *partially* ionizes in water is a **weak acid** or a **weak base**.

ANIMATION: Dissociation of Strong and Weak Acids



$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

FIGURE 13.6 Hydrochloric acid is a strong acid; it ionizes completely in water to form $H_3O^+(aq)$ and $Cl^-(aq)$.

TABLE 13.1 | Common Strong Acids

Formula	Name
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_3$	chloric acid
$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid (only one H^+ ionizes completely)

Strong Acids

ANIMATION: Acid Ionization

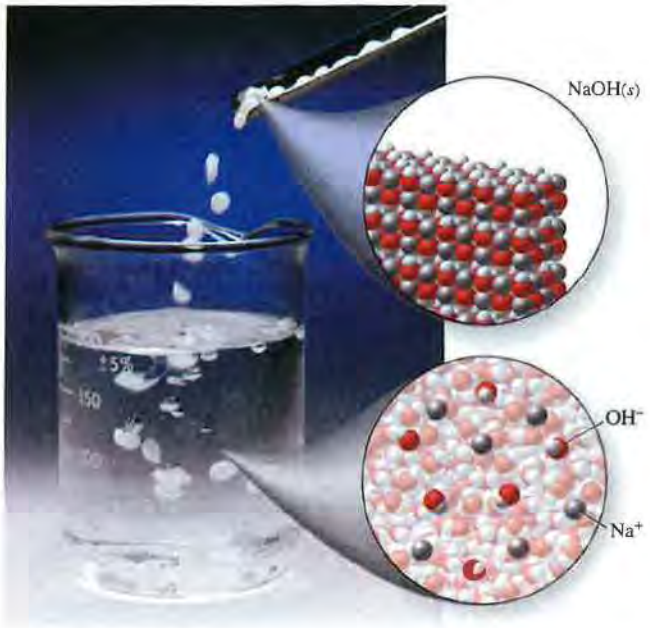
Strong acids ionize completely when dissolved in water. As shown in Figure 13.6, when the strong acid HCl ionizes in water, the resulting solution consists of only H_3O^+, Cl^-, and H_2O. It contains essentially no un-ionized HCl. Table 13.1 lists the common strong acids.

Some of the strong acids listed in Table 13.1 should be familiar to you. Hydrochloric acid, HCl, is the acid in your stomach that aids in digestion. It is also the acid that pool maintenance people like Jake use to wash the plaster of swimming pools. Sulfuric acid, H_2SO_4, is the acid in most car batteries, and it occurs in acid rain. Nitric acid, HNO_3, also occurs in acid rain.

Strong Bases

Strong bases dissociate completely when dissolved in water. As shown in Figure 13.7, the strong base sodium hydroxide, NaOH, dissociates completely in water to form $Na^+(aq)$ and $OH^-(aq)$. Essentially no undissociated $NaOH(aq)$ remains.

ANIMATION: Base Ionization



$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

FIGURE 13.7 The strong base NaOH dissociates completely in water to form $Na^+(aq)$ and $OH^-(aq)$.

Calcium oxide, CaO, has basic properties because, when added to water, it forms calcium hydroxide, $Ca(OH)_2(aq)$. Calcium oxide, commonly called lime, is used to reduce the acidity levels in soil. (When working with chemicals around your home or garden, it is safer to wear goggles and gloves.)



TABLE 13.2 | Common Strong Bases

Formula	Name	Formula*	Name
LiOH	lithium hydroxide	$Mg(OH)_2$	magnesium hydroxide
NaOH	sodium hydroxide	$Ca(OH)_2$	calcium hydroxide
KOH	potassium hydroxide	$Ba(OH)_2$	barium hydroxide

* Although the group IIA (2) metal hydroxides are not completely water soluble, they are strong bases because the amount that dissolves dissociates almost completely.

Most of the common strong bases are the ionic hydroxides of the group IA (1) and IIA (2) metals. Table 13.2 lists the common strong bases. Commercial oven and drain cleaners contain the strong base sodium hydroxide, NaOH. Sodium hydroxide is also used in processing paper and in canning olives. Calcium hydroxide, $Ca(OH)_2$, sometimes called slaked lime, forms when lime, CaO, mixes with water. Magnesium hydroxide, $Mg(OH)_2$, is a strong base used as an antacid, but as described in Chapter 11, it is relatively insoluble in water.

Weak Acids

How do we know if an acid is a weak acid? *Any acid that is not a strong acid is a weak acid.* A weak acid does not ionize completely when dissolved in water. An equilibrium forms between a weak acid and its conjugate base, so both the reactants and the products are present in solution. Examples of weak acids are citric acid, malic acid, and acetic acid. They are commonly found in fruits and other foods. Table 13.3 lists some weak acids and where they commonly occur.

TABLE 13.3 | Common Weak Acids

Formula	Name	Occurrence
CH_3CO_2H	acetic acid	vinegar, sour wine
H_2CO_3	carbonic acid	soda, blood
$H_3C_6H_5O_7$	citric acid	fruit, soda
HF	hydrofluoric acid	used in glass etching and semiconductor manufacturing
HOCl	hypochlorous acid	used to sanitize pools and drinking water
$HC_3H_5O_3$	lactic acid	milk
$HC_4H_4O_5$	malic acid	fruit
$H_2C_2O_4$	oxalic acid	nuts, cocoa, parsley, rhubarb
H_3PO_4	phosphoric acid	soda, blood
$H_2C_4H_4O_6$	tartaric acid	candy, wine, grapes



FIGURE 13.8 Acetic acid, CH_3CO_2H, is a weak acid because only a small portion of the acid molecules ionize in water. An aqueous solution of acetic acid has a much higher concentration of CH_3CO_2H molecules than of $CH_3CO_2^-(aq)$ and $H_3O^+(aq)$ ions. In this molecular-level image, there are 2 $CH_3CO_2^-(aq)$ ions and 10 $CH_3CO_2H(aq)$ molecules.

When a weak acid dissolves in water, usually fewer than 10% of the molecules transfer an H^+ to the water, leaving the conjugate base. The rest remain in the molecular form. The equilibrium lies far to the left. We use equilibrium arrows ($\rightleftharpoons$) when writing an equation to describe the ionization of a weak acid such as acetic acid, CH_3CO_2H, in water (Figure 13.8):

ANIMATION: Ionization of a Weak Acid

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons CH_3CO_2^-(aq) + H_3O^+(aq)$$

Other weak acids that ionize in water are cations that are the conjugate acids of weak bases. Ammonium ion, NH_4^+, is an example of a weak acid. It is found in the compound ammonium nitrate, NH_4NO_3, which Olivia uses to fertilize her plants. How does NH_4NO_3 act as an acid? Consider what happens when NH_4NO_3 dissolves in water. Like any other soluble ionic compound, it dissociates to its aqueous cations and anions:

$$NH_4NO_3(s) \longrightarrow NH_4^+(aq) + NO_3^-(aq)$$

Nitrate ion, NO_3^-, is the conjugate base of a very strong acid, so it does not behave as a base with water. The $NH_4^+(aq)$, however, reacts with water as a weak acid and donates an H^+ ion to water:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

Weak Bases

A weak base does not completely ionize when dissolved in water. An equilibrium forms between a weak base and its conjugate acid. Ammonia is a common weak base, as shown in Figure 13.9. Most common weak bases are organic compounds that contain the amine group ($-NH_2$), such as methylamine, CH_3NH_2. Methylamine is a weak base, so we show its ionization in water with equilibrium arrows:

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$$

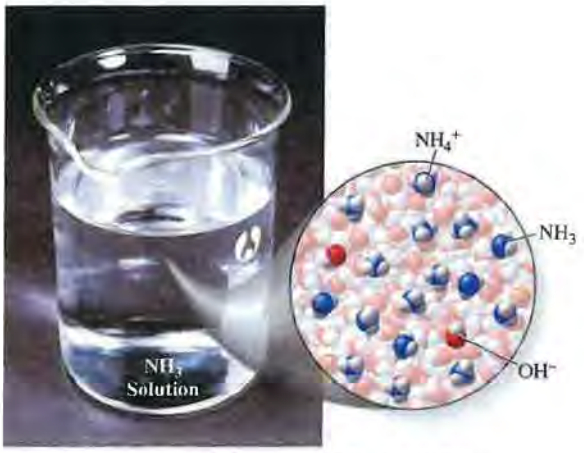
Other weak bases that ionize in water are anions that are the conjugate bases of weak acids. Hypochlorite ion, OCl^-, is an example of a weak base. It is found in the compound calcium hypochlorite, $Ca(OCl)_2$, that Jake uses to superchlorinate the water in swimming pools. How does $Ca(OCl)_2$ act as a base? Consider what happens when $Ca(OCl)_2$ dissolves in water. Like any other soluble ionic compound dissolved in water, it dissociates to its aqueous cations and anions:

$$Ca(OCl)_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2OCl^-(aq)$$

The $OCl^-(aq)$ reacts with water as a base and accepts an H^+ ion from water:

$$\underset{\text{weak base}}{OCl^-(aq)} + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$$

The anions that are the conjugate bases of weak acids are typically obtained from ionic compounds that contain the anion. For example, the hypochlorite ion used in pool sanitation can come from a variety of compounds, including $Ca(OCl)_2$. Some pool products contain sodium hypochlorite, NaOCl, instead. Identifying ionic compounds that behave as weak bases is relatively straightforward once you identify the anion. If the anion is the conjugate base of a weak acid, it will typically behave as



$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

FIGURE 13.9 Ammonia is a weak base, so only some of the NH_3 molecules ionize in water to form $NH_4^+(aq)$ and $OH^-(aq)$. An aqueous solution of ammonia has a much higher concentration of NH_3 molecules than of $NH_4^+(aq)$ and $OH^-(aq)$ ions.

TABLE 13.4 | Common Weak Bases

Formula	Name	Occurrence
NH_3	ammonia	glass cleaners
$CaCO_3$	calcium carbonate	antacids, minerals
$Ca(OCl)_2$	calcium hypochlorite	chlorine source for swimming pools
CH_3NH_2	methylamine	herring brine
$(CH_3)_3N$	trimethylamine	rotting fish

a base when added to water. For example, sodium acetate, $NaCH_3CO_2$, is a soluble ionic compound that dissociates to give $Na^+(aq)$ and $CH_3CO_2^-(aq)$ ions. Consulting Table 13.3, we see that acetic acid is a weak acid with the formula CH_3CO_2H. The acetate ion, $CH_3CO_2^-$, is the conjugate base of acetic acid, so we can predict that an ionic compound containing the acetate ion behaves as a weak base. On the other hand, if the anion is the conjugate base of a strong acid, it will not react as a base with water. For example, NaCl contains the conjugate base, Cl^-, of the strong acid HCl. Conjugate bases of strong acids do *not* react with water, so NaCl will *not* behave as a base.

Table 13.4 lists some common weak bases and where they are found. For those that are ionic compounds, identify the anion that acts as a base when the compound dissolves in water.

EXAMPLE 13.4 | Strong and Weak Acids and Bases in Aqueous Solution

Identify each of the following as a strong acid, weak acid, strong base, or weak base. Write an equation to describe its reaction with water.

(a) $CH_3CH_2NH_2(aq)$
(b) $HF(aq)$
(c) $NaF(aq)$
(d) $KOH(aq)$

Solution:

(a) The compound $CH_3CH_2NH_2$ is an organic compound called an amine because it contains the $-NH_2$ group. It is a weak base like NH_3. The equation for its ionization in water is

$$CH_3CH_2NH_2(aq) + H_2O(l) \rightleftharpoons CH_3CH_2NH_3^+(aq) + OH^-(aq)$$

Because $CH_3CH_2NH_2$ is a base, it ionizes in water to form $OH^-(aq)$ ions. Equilibrium arrows are used because $CH_3CH_2NH_2$ is a weak base and ionizes only partially.

(b) Hydrofluoric acid, HF, is an acid. Since it is not on the list of strong acids in Table 13.1, it is a weak acid. The equation for its ionization in water is

$$HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$$

Because HF is an acid, it reacts with water to form $H_3O^+(aq)$ ions. Equilibrium arrows are used because HF is a weak acid and ionizes only partially.

(c) Sodium fluoride is an ionic compound that completely dissociates in water to form $Na^+(aq)$ and $F^-(aq)$. The F^- ion is the conjugate base of the weak acid HF, so it is a weak base. The equation for its reaction with water is

$$F^-(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^-(aq)$$

Because F^- is a base, it reacts with water to form OH^- ions. Equilibrium arrows are used because F^- is a weak base.

(d) KOH is a strong base because it is a group IA (1) metal hydroxide. The equation for its dissociation in water is

$$KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$$

Because KOH is a base, it produces hydroxide ions in water. A one-directional arrow is used because KOH is a strong base and dissociates completely.

Practice Problem 13.4

Identify each of the following as a strong acid, weak acid, strong base, or weak base. Write an equation to describe its ionization in water.

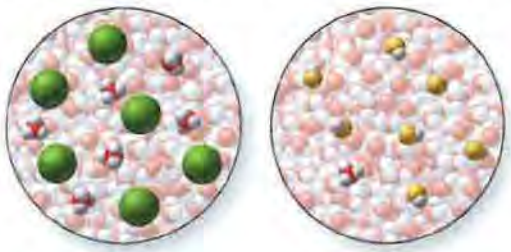
(a) $HI(aq)$
(b) $NaCH_3CO_2(aq)$
(c) $NH_4^+(aq)$
(d) $NH_3(aq)$

Further Practice: Questions 13.29 and 13.30 at the end of the chapter

Example 13.5 will help you distinguish between strong and weak acids at the molecular level.

EXAMPLE 13.5 | Molecular-Level Representations of Strong and Weak Acids

One of the diagrams shown represents a solution of HF, and the other represents an aqueous solution of HCl. Which is which? Explain your reasoning.



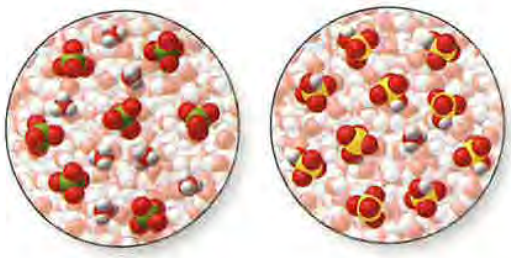
A B

Solution:

The difference between the two diagrams is the degree of ionization. Diagram A shows no acid molecules, indicating complete ionization of the acid. It represents a strong acid. Diagram B shows many un-ionized acid molecules and a few that are ionized, indicating partial ionization. It represents a weak acid. Of the two acids, HCl is the strong acid and is represented by diagram A. Diagram B represents a solution of HF, the weak acid.

Practice Problem 13.5

One of the diagrams represents a solution of $HClO_4$, and the other represents an aqueous solution of HSO_4^-. Which is which? Explain your reasoning.



A B

Further Practice: Questions 13.31 and 13.32 at the end of the chapter

13.3 | RELATIVE STRENGTHS OF WEAK ACIDS

Some of the sour candies made at the factory where Sara works are coated with a mixture of sugar and citric acid. Citric acid has a stronger sour flavor than many other edible acids. Why? Is it more acidic than other weak acids? How can we compare the acid strength of different weak acids? Acid strength depends on the relative number of acid molecules that ionize when dissolved in water—that is, the degree of ionization. The percentage of molecules that ionize varies among weak acids. The stronger the acid, the larger its equilibrium constant value for ionization in water, and the greater the percentage of $H_3O^+(aq)$ and conjugate base ions produced. Remember from Chapter 12 that an equilibrium constant K_{eq} describes the relative amounts of reactants and products. The larger the value of K_{eq}, the more products exist at equilibrium.

Consider the ionization in water of two weak acids: acetic acid, CH_3CO_2H, and hydrocyanic acid, HCN:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq) \qquad K_{eq} = 1.8 \times 10^{-5}$$

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq) \qquad K_{eq} = 6.2 \times 10^{-10}$$

Which acid, CH_3CO_2H or HCN, is stronger? Their equilibrium constants provide the answer. The stronger acid ionizes to a larger extent, producing more $H_3O^+(aq)$ ions and conjugate base than the weaker acid. The stronger acid has more products at equilibrium and therefore has an equilibrium that lies farther to the right than the weaker acid. Because acetic acid, CH_3CO_2H, has the greater equilibrium constant value for ionization, it is the stronger acid.

The equilibrium constant represents the ratio of the concentration of products to reactants at equilibrium, leaving out any substances that are in the liquid or solid phase. For example, the equilibrium constant expression for the ionization of acetic acid, CH_3CO_2H, in water is

$$K_{eq} = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

Recall from Chapter 12 that by convention equilibrium constants are written without units.

Acid Ionization Constants

To distinguish the equilibrium constant for the ionization of acids in water from other reactions acids can undergo, chemists use a special symbol, K_a, the **acid ionization constant**. It describes the equilibrium that forms when an acid reacts with water. Thus, the larger the K_a value, the stronger the acid. The K_{eq} values given for the ionization of acetic acid and hydrocyanic acid are also their K_a values:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq) \qquad K_a = 1.8 \times 10^{-5}$$

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq) \qquad K_a = 6.2 \times 10^{-10}$$

TABLE 13.5 | Weak Acids, Their K_a Values at 25°C, and Their Conjugate Bases

	Acid	K_a Values	Conjugate Base	
Strongest acids	HF	6.3×10^{-4}	F^-	Weakest bases
↑	HNO_2	5.6×10^{-4}	NO_2^-	↓
	HCO_2H	1.8×10^{-4}	HCO_2^-	
	CH_3CO_2H	1.8×10^{-5}	$CH_3CO_2^-$	
	HOCl	4.0×10^{-8}	OCl^-	
	HCN	6.2×10^{-10}	CN^-	
Weakest acids	NH_4^+	5.6×10^{-10}	NH_3	Strongest bases

Given K_a values for weak acids, we can compare the relative amounts of $H_3O^+(aq)$ in different weak acid solutions if the solutions are identical in concentration of acid. What can we say about the relative amounts of $H_3O^+(aq)$ in a 0.10 M CH_3CO_2H solution compared to a 0.10 M HCN solution? Acetic acid, CH_3CO_2H, with the larger K_a value, ionizes to the larger extent, yielding a greater concentration of H_3O^+. However, the concentration of H_3O^+ in both weak acid solutions is much less than the H_3O^+ concentration in a strong acid solution of the same concentration, such as 0.10 M HCl. We cannot easily compare H_3O^+ concentrations in solutions of weak acids of different concentrations because the concentration of the acid is also a factor. Table 13.5 lists several weak acids and their K_a values, along with the conjugate base of each acid. It also shows that the strength of the conjugate base is inversely related to the strength of the acid. In other words, the stronger the acid, the weaker the conjugate base. The conjugate bases of strong acids (such as Cl^-) are so weak that they do not react with water.

How is acid strength different from acid concentration? For example, when comparing 0.1 M HCl and 0.2 M HF, which is the stronger acid? Which is the more concentrated acid?

EXAMPLE 13.6 | Significance of K_a

Which acid ionizes in water to the greater extent, hypochlorous acid, HOCl, or hydrocyanic acid, HCN? Consult Table 13.5 for K_a values.

Solution:

From Table 13.5, we see that the K_a value for HOCl, 4.0×10^{-8}, is larger than the K_a value for HCN, 6.2×10^{-10}. The larger K_a value for HOCl indicates that a greater percentage of HOCl reacts, so it ionizes to H_3O^+ and OCl^- to a greater extent.

Practice Problem 13.6

Which solution has the greater concentration of H_3O^+, a 0.10 M solution of hypochlorous acid, HOCl, or a 0.10 M solution of hydrocyanic acid, HCN?

Further Practice: Questions 13.37 and 13.38 at the end of the chapter



FIGURE 13.10 Polyprotic acids have more than one acidic hydrogen. How many acidic hydrogen atoms does each of these acids have?

Polyprotic Acids

Some acids contain more than one acidic hydrogen atom. Such an acid, which can donate more than one H^+ ion, is called a **polyprotic acid** (Figure 13.10). For example, a phosphoric acid molecule, H_3PO_4, has three acidic hydrogens, each

bonded to an oxygen atom. Carbonic acid (H_2CO_3) and sulfuric acid (H_2SO_4) are also polyprotic acids. Polyprotic acids do not lose all of their acidic hydrogen atoms in water to the same extent. We consider the loss of hydrogens by a polyprotic acid as a stepwise process. For example, consider carbonic acid, which we treat as the dissolved form of CO_2:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

When the H_2CO_3 ionizes in water, only small amounts of $HCO_3^-(aq)$ and $H_3O^+(aq)$ are produced. Significantly less $CO_3^{2-}(aq)$ is present in the solution. Two acid ionization equations are used to describe the equilibria that exist in solution, each with its own K_a value:

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq) \qquad K_{a1} = 4.5 \times 10^{-7}$$
$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq) \qquad K_{a2} = 4.7 \times 10^{-11}$$

Sulfuric acid, H_2SO_4, is a polyprotic acid that ionizes completely in water to form $HSO_4^-(aq)$ and $H_3O^+(aq)$:

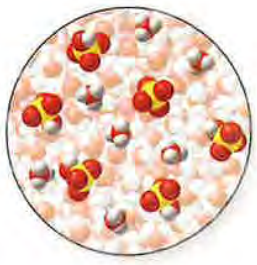
$$H_2SO_4(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + HSO_4^-(aq) \qquad \text{complete ionization}$$

Once $HSO_4^-(aq)$ forms, it also acts as an acid, but as a weak acid:

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq) \qquad K_a = 1.2 \times 10^{-2}$$

The low K_a value for the second ionization shows that the equilibrium lies toward the reactants; only a small portion of the $HSO_4^-(aq)$ ionizes. A smaller portion of the $HSO_4^-(aq)$ ionizes compared to H_2SO_4 because it is more difficult to remove a positively charged H^+ ion from a negatively charged HSO_4^- ion than from a neutral H_2SO_4 molecule. An aqueous solution of sulfuric acid, then, contains essentially no $H_2SO_4(aq)$, a relatively large amount of $HSO_4^-(aq)$, and a small amount of $SO_4^{2-}(aq)$ (Figure 13.11).

The K_a values are often labeled to indicate the particular step in the overall ionization process (K_{a1}, K_{a2}, K_{a3}, etc.). Table 13.6 lists several polyprotic acids and their K_a values. Notice that for successive ionizations, the K_a values decrease significantly as H^+ ions are removed.



$H_2SO_4(aq)$

FIGURE 13.11 Sulfuric acid, H_2SO_4, is a strong polyprotic acid. It ionizes completely in water to form $HSO_4^-(aq)$ and $H_3O^+(aq)$. The $HSO_4^-(aq)$ produced also ionizes, but to a much lesser extent. In an aqueous solution of H_2SO_4, there are essentially no H_2SO_4 molecules, lots of $HSO_4^-(aq)$ ions, and a small fraction of $SO_4^{2-}(aq)$ ions. Can you identify the $HSO_4^-(aq)$ and $SO_4^{2-}(aq)$ ions in this figure?

TABLE 13.6 | Acid Ionization Constants for Some Polyprotic Acids (at 25°C)

Formula	Name	K_{a1}	K_{a2}	K_{a3}
H_2CO_3	carbonic acid	4.5×10^{-7}	4.7×10^{-11}	
$H_3C_6H_5O_7$	citric acid	7.4×10^{-4}	1.7×10^{-5}	4.0×10^{-7}
H_2S	hydrosulfuric acid	8.9×10^{-8}	1.0×10^{-19}	
$H_2C_2O_4$	oxalic acid	5.6×10^{-2}	1.5×10^{-4}	
H_3PO_4	phosphoric acid	6.9×10^{-3}	6.2×10^{-8}	4.8×10^{-13}
H_2SO_4	sulfuric acid	strong	1.2×10^{-2}	
$H_2C_4H_4O_6$	tartaric acid	1.0×10^{-3}	4.3×10^{-5}	

EXAMPLE 13.7 | Behavior of Polyprotic Acids in Water

The tartaric acid, $H_2C_4H_4O_6$, in grapes promotes a crisp flavor and graceful aging in wine. Its K_a values are given in Table 13.6.

(a) Write equations that show the ionization of tartaric acid in water.
(b) Besides water, which ion or molecule has the highest concentration in solution when tartaric acid is added to water?

Solution:

(a) Tartaric acid has two K_a values, so it has two acidic hydrogen atoms. Two equations are required to describe its behavior in water. The first shows the ionization of the first H^+. The K_a that corresponds to this process is the largest K_a value because the first H^+ always ionizes to the greater extent.

$$H_2C_4H_4O_6(aq) + H_2O(l) \rightleftharpoons HC_4H_4O_6^-(aq) + H_3O^+(aq) \quad K_{a1} = 1.0 \times 10^{-3}$$
$$HC_4H_4O_6^-(aq) + H_2O(l) \rightleftharpoons C_4H_4O_6^{2-}(aq) + H_3O^+(aq) \quad K_{a2} = 4.3 \times 10^{-5}$$

(b) The K_{a1} value for $H_2C_4H_4O_6$ shows that it is a weak acid and ionizes only to a small extent. That means that *most* of the $H_2C_4H_4O_6$ does not ionize and is present in water in the highest concentration.

Practice Problem 13.7

Oxalic acid, $H_2C_2O_4$, occurs in plants and foods such as parsley, rhubarb, almonds, and green beans. Its K_a values are listed in Table 13.6.

(a) Write equations that show the ionization of oxalic acid in water.
(b) Besides water, which ion or molecule has the highest concentration in solution when oxalic acid is added to water?

Further Practice: Questions 13.45 and 13.46 at the end of the chapter

13.4 | ACIDIC, BASIC, AND NEUTRAL SOLUTIONS

When Olivia tests a soil solution to make sure it is not too acidic, what specifically is she testing for? Remember that acids ionize when dissolved in water to form hydronium ions, $H_3O^+(aq)$. Does the presence of H_3O^+ ions in solution indicate an acidic solution? No. All aqueous solutions, and even pure water, have some H_3O^+ ions (and OH^- ions) in solution. It is the *relative amounts* of these ions that make the difference. In any **acidic solution**, the H_3O^+ ion concentration is greater than the OH^- ion concentration. In any **basic solution**, the OH^- ion concentration is greater than the H_3O^+ ion concentration. A **neutral solution** is neither acidic nor basic, but has equal concentrations of H_3O^+ and OH^- ions.

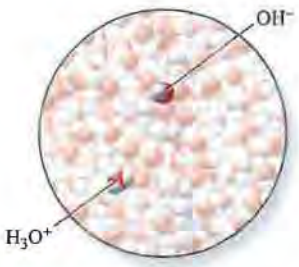


FIGURE 13.12 To a very small extent, water reacts with itself by transferring an H^+ ion from one water molecule to another, producing $H_3O^+(aq)$ and $OH^-(aq)$ ions.

The Ion-Product Constant of Water

Pure water, H_2O, is amphoteric. It acts as *both* an acid and a base, but only to a small extent. Water reacts with itself by a process called **self-ionization** (Figure 13.12), in which an H^+ ion is transferred from one water molecule to another:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Self-ionization is an equilibrium process, so we can write an equilibrium constant expression for it:

$$K_{eq} = [H_3O^+][OH^-]$$

Because the reactant, H_2O, is a liquid, it is not included in the equilibrium constant expression. This equilibrium constant, called the **ion-product constant of water**, is so widely used that it is given a special symbol, K_w. At 25°C, K_w has a value of 1.0×10^{-14}:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

The small value of this equilibrium constant indicates that the equilibrium lies very far to the left, so the concentrations of H_3O^+ and OH^- are very small. In pure water the concentrations of these ions must be equal, because they are produced in a 1:1 ratio:

$$[H_3O^+] = [OH^-]$$

From the K_w expression, we know that these concentrations must multiply together to give the K_w value of 1.0×10^{-14}:

$$1.0 \times 10^{-14} = [H_3O^+][OH^-]$$

For these concentrations to be equal, each must have a concentration of 1.0×10^{-7} *M*, because if we substitute this value for the H_3O^+ concentration and the OH^- concentration,

$$1.0 \times 10^{-14} = (1.0 \times 10^{-7}\ M) \times (1.0 \times 10^{-7}\ M)$$

we get

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ M$$

For our purposes, we will assume the temperature to be 25°C, so we will not have to worry about using K_w values at other temperatures.

In a neutral solution such as pure water, the H_3O^+ ion and OH^- ion concentrations are equal. If an acid or base is added to water, the concentrations of both ions change. However, their product continues to be 1.0×10^{-14}, the value of K_w that is constant. How can this be? The reaction of an acid with water increases the H_3O^+ ion concentration to a value greater than 1.0×10^{-7} *M*. At the same time, the OH^- ion concentration decreases, with the result that the ion-product constant for water remains equal to 1.0×10^{-14}. In a similar fashion, the addition of a base to water increases the OH^- ion concentration and decreases the H_3O^+ ion concentration, but the ion-product constant remains unchanged. Table 13.7 summarizes the definitions of neutral, acidic, and basic solutions in aqueous solution.

Remember from Chapter 12 that square brackets around a formula represent the molar concentration of that substance. For example, $[H_3O^+]$ represents the molar concentration of hydronium ion.

Calculating H_3O^+ and OH^- Ion Concentrations

If we know either the H_3O^+ ion concentration or the OH^- ion concentration, we can calculate the other unknown concentration using the ion-product constant of water and its expression:

$$K_w = [H_3O^+][OH^-]$$

TABLE 13.7 | Definitions of Neutral, Acidic, and Basic Solutions in Aqueous Solution

Type of Solution	Relative Concentrations	$[H_3O^+]$	$[OH^-]$	K_w
neutral	$[H_3O^+] = [OH^-]$	$= 1.0 \times 10^{-7}\ M$	$= 1.0 \times 10^{-7}\ M$	1.0×10^{-14}
acidic	$[H_3O^+] > [OH^-]$	$> 1.0 \times 10^{-7}\ M$	$< 1.0 \times 10^{-7}\ M$	1.0×10^{-14}
basic	$[OH^-] > [H_3O^+]$	$< 1.0 \times 10^{-7}\ M$	$> 1.0 \times 10^{-7}\ M$	1.0×10^{-14}

If we know the H_3O^+ ion concentration, $[H_3O^+]$, we rearrange the K_w expression to solve for the OH^- ion concentration, $[OH^-]$:

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$

If we know the OH^- ion concentration, $[OH^-]$, we rearrange the K_w expression to solve for the H_3O^+ ion concentration, $[H_3O^+]$:

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{[OH^-]}$$

These calculations are shown in Example 13.8.

EXAMPLE 13.8 | Calculating H_3O^+ and OH^- Ion Concentrations

Given the concentration of H_3O^+ ion in water, calculate the concentration of OH^- ion in the solution. Identify each solution as acidic, basic, or neutral:

(a) $[H_3O^+] = 1.0 \times 10^{-9}\ M$
(b) $[H_3O^+] = 0.0010\ M$

Solution:

In any aqueous solution at 25°C, the product of the H_3O^+ ion concentration and the OH^- concentration equals 1.0×10^{-14}:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

We rearrange the equation to solve for hydroxide ion concentration:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$

(a) We substitute the H_3O^+ ion concentration of $1.0 \times 10^{-9}\ M$ into the equation to solve for the OH^- ion concentration. Recall that equilibrium constants do not have units. When you analyze the units in calculations that involve equilibrium constants, it might appear that they will not cancel properly. Assume that if you're calculating the concentration of a chemical species involved in the equilibrium that it will have units of molarity.

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-9}\ M} = 1.0 \times 10^{-5} M$$

The solution is basic because $[OH^-] > [H_3O^+]$ and because $[H_3O^+] < 1.0 \times 10^{-7}\ M$.

(b) We substitute the H_3O^+ ion concentration of $0.0010\ M$ into the equation to solve for the OH^- ion concentration:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.0010\ M} = 1.0 \times 10^{-11} M$$

The solution is acidic because $[H_3O^+] > [OH^-]$ and because $[H_3O^+] > 1.0 \times 10^{-7}\ M$.

Practice Problem 13.8

Given the concentration of OH^- ion in each solution, calculate the concentration of H_3O^+ in the solution. Identify each solution as acidic, basic, or neutral.

(a) $[OH^-] = 1.0 \times 10^{-8}\ M$
(b) $[OH^-] = 0.010\ M$

Further Practice: Questions 13.55 and 13.56 at the end of the chapter

Suppose we are using a solution in a laboratory that is labeled 0.010 *M* HCl. Can we determine its concentrations of H_3O^+ ion and OH^- ion? Hydrochloric acid, HCl, is a strong acid, so it ionizes completely to form H_3O^+ ions and Cl^- ions. The concentration of H_3O^+ is actually the concentration of HCl labeled on the bottle:

$$HCl(aq) + H_2O(l) \longrightarrow \underset{0.010\ M}{H_3O^+(aq)} + \underset{0.010\ M}{Cl^-(aq)}$$

From the H_3O^+ concentration, $[H_3O^+]$, and the value of K_w, we can determine the OH^- ion concentration, $[OH^-]$, as we did in Example 13.8. Work the problem and assure yourself that the value is $1.0 \times 10^{-12}\ M$ OH^-.

Suppose we have a 0.10 *M* solution of sodium hydroxide, NaOH, in the laboratory. What are the H_3O^+ and OH^- ion concentrations in this solution? Sodium hydroxide is a base, so it forms $OH^-(aq)$ ions when dissolved in water. Because it is a strong base with only one OH^- group, $[OH^-]$ is equal to the concentration of NaOH labeled on the bottle:

$$NaOH(s) \xrightarrow{H_2O} \underset{0.10\ M}{Na^+(aq)} + \underset{0.10\ M}{OH^-(aq)}$$

The hydroxide ion concentration, $[OH^-]$, is 0.10 *M*. Using the K_w expression calculate the H_3O^+ ion concentration, $[H_3O^+]$, as $1.0 \times 10^{-13}\ M$.

When strong bases such as $Ba(OH)_2$ contain more than one OH^- ion per formula unit, the OH^- concentration is twice the concentration of the $Ba(OH)_2$ solution. Why?

EXAMPLE 13.9 | Calculating $[H_3O^+]$ and $[OH^-]$ in Strong Acid and Strong Base Solutions

Stomach acid is about 1.0 *M* HCl. What is the H_3O^+ concentration in stomach acid? What is the OH^- concentration?

Solution:

Hydrochloric acid, HCl, is a strong acid, so it ionizes completely in water. For every mole of HCl that ionizes, one mole of H_3O^+ forms. The concentration of H_3O^+, then, is 1.0 *M*. The OH^- concentration is determined from the H_3O^+ concentration and the expression for the ion-product constant for water:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$

$$= \frac{1.0 \times 10^{-14}}{1.0\ M}$$

$$= 1.0 \times 10^{-14}\ M$$

Practice Problem 13.9

Sodium hydroxide is sometimes called lye or caustic soda. What is the OH^- concentration in a 0.85 *M* lye solution? What is the H_3O^+ concentration?

Further Practice: Questions 13.57 and 13.58 at the end of the chapter

Now suppose that we are using a 0.010 *M* acetic acid solution, $CH_3CO_2H(aq)$, in the lab. Recall that a 0.010 *M* HCl solution has an H_3O^+ ion concentration of 0.010 *M* because HCl is a strong acid—it completely ionizes. The H_3O^+ concentration in a 0.010 *M* CH_3CO_2H solution is only 4.2×10^{-4} *M*. Why? Acetic acid is a weak acid that does not completely ionize. In a 0.010 *M* solution, only about 4.2% of the acetic acid molecules ionize, while in a hydrochloric acid solution, 100% of the HCl molecules ionize.

13.5 | THE pH SCALE

Jake tests the acidity of pool water using pH indicators and a pH meter. Olivia also measures the pH of soil to determine whether to add lime. Perhaps you have heard the term pH, as it is used frequently in advertisements for shampoos and skin-care products. The pH of solutions is easily measured and tells us about the acidity or basicity of a substance. In this section, we will explore how pH is related to the concentration of H_3O^+ (and OH^-) and examine the significance of the pH scale.

Calculating pH

Concentrations of H_3O^+ ions can vary widely among concentrated and dilute acids, so it is convenient to express the acidity of aqueous solutions on a *pH scale*. The **pH** of a solution is defined as the negative logarithm (base 10) of the H_3O^+ concentration:

$$\text{pH} = -\log[H_3O^+]$$

For example, the H_3O^+ concentration in pure water is 1.0×10^{-7} *M*. Substituting this value into the equation for pH gives a pH of 7.00:

$$\begin{aligned}\text{pH} &= -\log[H_3O^+] \\ &= -\log(1.0 \times 10^{-7}) \\ &= 7.00\end{aligned}$$



The log function is a button on most calculators. See Math Toolbox 13.1 at the end of the chapter for instructions on using this function on your calculator.

Like pure water, all neutral solutions have equal concentrations of H_3O^+ and OH^- (1.0×10^{-7} *M* each), so their pH values are 7. Acidic solutions have pH values less than 7, and basic solutions have pH values greater than 7.

To better understand the meaning of the log function, consider the pH values of solutions that range in H_3O^+ concentration from 1.0×10^{-1} *M* to 1.0×10^{-14} *M*, as shown in Figure 13.13. Notice that each pH value equals the exponent in the H_3O^+ concentration value, but with a positive sign. Use the logarithmic definition of pH to convince yourself that the pH values listed there can be calculated from the H_3O^+ concentration values given.

How do we apply rules for significant figures to pH calculations? The rule is that the number of decimal places in a pH value should equal the number of significant figures in the H_3O^+ concentration value.

The relationship between pH and concentration allows us to determine the pH without a calculator only when the H_3O^+ concentration is exactly a power of 10. When it is not—for example, 5.0×10^{-4} *M*—you can use the log function on your calculator to determine the pH:

$$\begin{aligned}\text{pH} &= -\log[H_3O^+] \\ &= -\log(5.0 \times 10^{-4}) \\ &= 3.30\end{aligned}$$

The pH of a strong acid solution can be determined if you know the concentration of the acid. For the strong acid HNO_3, for example, the concentration of H_3O^+

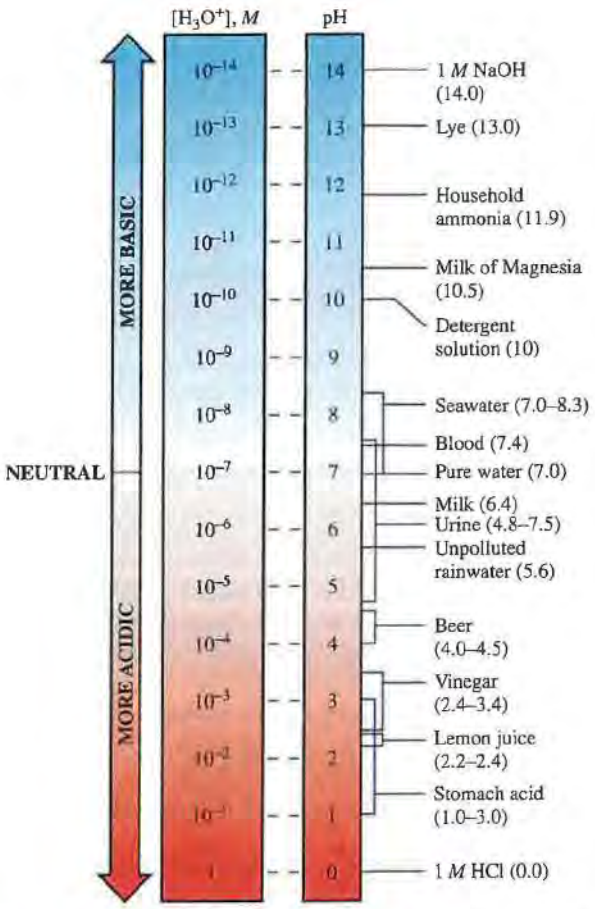


FIGURE 13.13 Aqueous solutions we use in our daily lives generally have pH values that range from 0 to 14. However, some acids and bases used in the laboratory are sometimes so concentrated that their pH values are less than 0 or greater than 14. What concentrations of HCl would give a pH less than zero? What concentrations of NaOH would give a pH greater than 14?

equals the HNO_3 concentration because strong acids ionize completely. What is the pH of a 0.10 *M* HNO_3 solution? Because HNO_3 is a strong acid, the concentration of H_3O^+ ion in the solution is 0.10 *M*, or 1.0×10^{-1} *M*. The pH is calculated by substituting the H_3O^+ concentration into the following equation:

$$\begin{aligned} \text{pH} &= -\log[H_3O^+] \\ &= -\log(1.0 \times 10^{-1}) \\ &= 1.00 \end{aligned}$$

We can also determine the pH of a strong base solution, such as a 0.010 *M* NaOH solution. Strong bases ionize completely, so the concentration of OH^- ion is the same as the concentration of NaOH, 0.010 *M* or 1.0×10^{-2} *M*. Because the pH equation requires the H_3O^+ concentration, we first calculate the H_3O^+ concentration from the OH^- concentration using the K_w expression:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.010\ M} = 1.0 \times 10^{-12} M$$

Now that we know the H_3O^+ concentration, 1.0×10^{-12} *M*, we can calculate the pH of the solution:

$$\begin{aligned} pH &= -\log[H_3O^+] \\ &= -\log(1.0 \times 10^{-12}) \\ &= 12.00 \end{aligned}$$

EXAMPLE 13.10 | Calculating pH

What is the pH of each of the following solutions? Once you've done the calculation, check your answer to make sure it makes sense.

(a) 0.0010 *M* HBr
(b) 0.035 *M* HNO_3
(c) 0.035 *M* KOH

Solution:

(a) Hydrobromic acid, HBr(*aq*), is a strong acid, which ionizes completely, so the H_3O^+ concentration equals the concentration of HBr:

$$[H_3O^+] = 0.0010\ M$$



The pH can be calculated from the H_3O^+ concentration:

$$\begin{aligned} pH &= -\log[H_3O^+] \\ &= -\log(0.0010) \\ &= 3.00 \end{aligned}$$

A pH of 3.00 makes sense for an acid because it is less than 7, as we would expect for a solution of any acid.

(b) Nitric acid, $HNO_3(aq)$, is also a strong acid, so the H_3O^+ concentration equals the concentration of HNO_3:

$$[H_3O^+] = 0.035\ M$$

The pH can be calculated from the H_3O^+ concentration:

$$\begin{aligned} pH &= -\log[H_3O^+] \\ &= -\log(0.035) \\ &= 1.46 \end{aligned}$$

We would expect a lower pH for HNO_3 than for HBr because the HNO_3 solution was more concentrated.

(c) Potassium hydroxide, KOH(*aq*), is a strong base, which dissociates completely to form $K^+(aq)$ and $OH^-(aq)$ ions. The concentration of KOH, therefore, is also the OH^- concentration:

$$[OH^-] = 0.035\ M$$

In order to calculate pH, we need to determine the H_3O^+ concentration from the OH^- concentration and the K_w expression:

$$[H_3O^+][OH^-] = K_w = 1.0 \times 10^{-14}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.035\ M} = 2.9 \times 10^{-13} M$$

Now that we know the H_3O^+ concentration, we can calculate the pH of the solution:

$$\begin{aligned} \text{pH} &= -\log[H_3O^+] \\ &= -\log(2.9 \times 10^{-13}) \\ &= 12.54 \end{aligned}$$

A high pH (greater than 7) is expected because this is a solution of a base.

Practice Problem 13.10

What is the pH of each of the following solutions? Once you've done the calculation, check your answer to make sure it makes sense.

(a) 0.00085 *M* HCl
(b) 0.10 *M* NaOH
(c) 1.0 *M* HNO_3

Further Practice: Questions 13.71 and 13.72 at the end of the chapter

In very dilute solutions of strong acids and bases, determining the pH is more complicated. For example, consider a 1.0×10^{-8} *M* HCl solution. As a strong acid, we expect HCl to completely ionize, but an H_3O^+ concentration of 1.0×10^{-8} *M* does not make sense because acidic solutions have H_3O^+ concentrations greater than the 1.0×10^{-7} *M* H_3O^+ from pure water. When you encounter very dilute solutions of acids with concentrations less than 1.0×10^{-7} *M*, assume that the pH is slightly less than 7.0. If you have a solution of a base with a concentration less than 1.0×10^{-7} *M*, assume that the pH is slightly greater than 7.0.

Calculating pOH

Just as pH can be calculated from the H_3O^+ concentration, a value called pOH can be calculated from the OH^- concentration of the solution. The pOH of a solution is defined as the negative logarithm (base 10) of the hydroxide ion concentration:

$$\text{pOH} = -\log[OH^-]$$

Earlier in this section we found that the pH of a 0.010 *M* NaOH solution is 12.00. The pOH of this strong base can be calculated directly, because the OH^- concentration equals the concentration of NaOH, 0.010 *M*. This value can be substituted into the pOH equation:

$$\begin{aligned} \text{pOH} &= -\log[OH^-] \\ &= -\log(0.010) \\ &= -\log(1.0 \times 10^{-2}) \\ &= 2.00 \end{aligned}$$

Thus, the pOH of this 0.010 *M* NaOH solution is 2.00. Note that the pH and pOH values total 14.00. This is true for any aqueous solution at 25°C.

$$\text{pH} + \text{pOH} = 14.00$$

This relationship allows us to determine the pH or pOH of a solution if one value or the other is known. For example, if the pOH of a solution is 4.00, its pH is calculated by subtracting the pOH from 14.00:

$$\begin{aligned} \text{pH} + \text{pOH} &= 14.00 \\ \text{pH} &= 14.00 - \text{pOH} \\ &= 14.00 - 4.00 \\ &= 10.00 \end{aligned}$$

Why does pH + pOH = 14.00? The relationship can be derived mathematically from the K_w expression of water:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Taking the negative log of both sides of the equation gives

$$-\log([H_3O^+][OH^-]) = -\log(1.0 \times 10^{-14})$$

The log of a product of two values equals the *sum* of the logs of the terms:

$$-\log[H_3O^+] - \log[OH^-] = -\log(1.0 \times 10^{-14})$$

We know $-\log[H_3O^+]$ and $-\log[OH^-]$ equal pH and pOH, respectively, so we can substitute pH and pOH into this equation:

$$\text{pH} + \text{pOH} = 14.00$$



A



B

FIGURE 13.14 The color of hydrangeas depends on soil pH and the presence of aluminum ion. (A) In acidic soil in the pH range of 5.5 to 6.5, and in the presence of adequate aluminum, flowers appear deep blue. (B) In more alkaline soil (pH 7.0 to 7.5), the flowers are pink.

Calculating Concentrations from pH or pOH

Olivia has beautiful blue hydrangeas in her garden, but she wants some of her plants to show flowers that are pink (Figure 13.14). After some research, she discovers that the flower color, in part, depends on the pH of the soil. Blue flowers occur when the soil is in the pH range of 5.5 to 6.5. For pink flowers, the pH needs to be slightly more alkaline, in the range of 7.0 to 7.5. When she measures the pH of a soil solution from her garden, she learns that it is slightly acidic with a pH of 6.20. She learns that making the soil more alkaline to produce pink flowers can be achieved by adding calcium oxide (lime) to the soil. Her endeavor into soil pH has Olivia curious about the concentration of H_3O^+ ion in the soil solution. She needs an equation that relates $[H_3O^+]$ to pH, and the pH equation accomplishes this:

$$pH = -\log[H_3O^+]$$

To get $[H_3O^+]$, we must rearrange the equation so that $[H_3O^+]$ is on one side by itself. First, we multiply both sides by −1:

$$-pH = \log[H_3O^+]$$

To get rid of the log expression in front of $[H_3O^+]$, we take the inverse log of both sides. The H_3O^+ concentration, then, equals the inverse log of the negative pH:

$$\text{Inverse log } (-pH) = [H_3O^+]$$

Taking the inverse log of a value (x) is the same as 10 raised to the x power. In other words, *the H_3O^+ concentration equals 10 raised to the –pH power:*

$$10^{-pH} = [H_3O^+]$$

The H_3O^+ concentration in a soil sample with a pH of 6.20 can be calculated using this equation:

$$[H_3O^+] = 10^{-pH} = 10^{-6.20} = 6.3 \times 10^{-7}\ M$$

The OH^- concentration can be determined from the pOH with a similar equation:

$$[OH^-] = 10^{-pOH}$$

The relationships among pH, pOH, $[H_3O^+]$, and $[OH^-]$ are represented in Figure 13.15. If you know one of these values for a solution, you can calculate any of the other three. For example, consider the values shown in Figure 13.16. Notice how the four quantities $[H_3O^+]$, pH, pOH, and $[OH^-]$ are related. Note that there is more than one way to convert from one value to another, as illustrated in Example 13.11.

The inverse log function on your calculator is the 10^x function, which is normally found above the log function as an alternate function for the same key. For more information on how to use the inverse log function, see Math Toolbox 13.1.

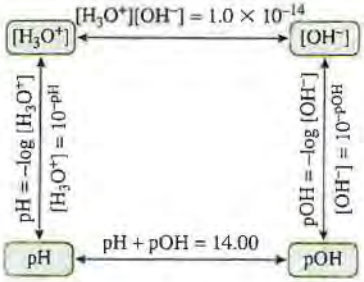


FIGURE 13.15 If you know pH, pOH, $[H_3O^+]$, or $[OH^-]$, you can calculate any of the other quantities.

EXAMPLE 13.11 | Converting Between pH, pOH, $[H_3O^+]$, and $[OH^-]$

Jake measured the pH of water in a swimming pool as 8.10. What is the OH^- concentration in the pool water?

Solution:

Figure 13.15 shows two alternate routes we can take for this calculation. One route is first to convert pH to pOH and then convert pOH to the OH^- concentration:



To convert from pH to pOH, we use the relationship between them:

$$pH + pOH = 14.00$$
$$8.10 + pOH = 14.00$$
$$pOH = 5.90$$

To convert from pOH to $[OH^-]$, we use the relationship that allows us to solve for hydroxide ion concentration:

$$[OH^-] = 10^{-pOH} = 10^{-5.90} = 1.3 \times 10^{-6}\ M$$

MATH TOOLBOX 13.1

The stepwise calculation is summarized as follows:

pH → (pH + pOH = 14.00) → pOH → ($[OH^-] = 10^{-pOH}$) → $[OH^-]$

An alternative route would be to convert first from pH to $[H_3O^+]$ and then from $[H_3O^+]$ to $[OH^-]$:

pH → $[H_3O^+]$ → $[OH^-]$

Try this route. You should get the same answer.

Practice Problem 13.11

When the pH of water in a lake falls below about 4.5, the lake may be considered dead because few organisms can survive in such an acidic environment. What is the pH of a lake that has an OH^- concentration of $1.0 \times 10^{-9}\ M$? Would this lake be considered dead?

Further Practice: Questions 13.73 and 13.74 at the end of the chapter

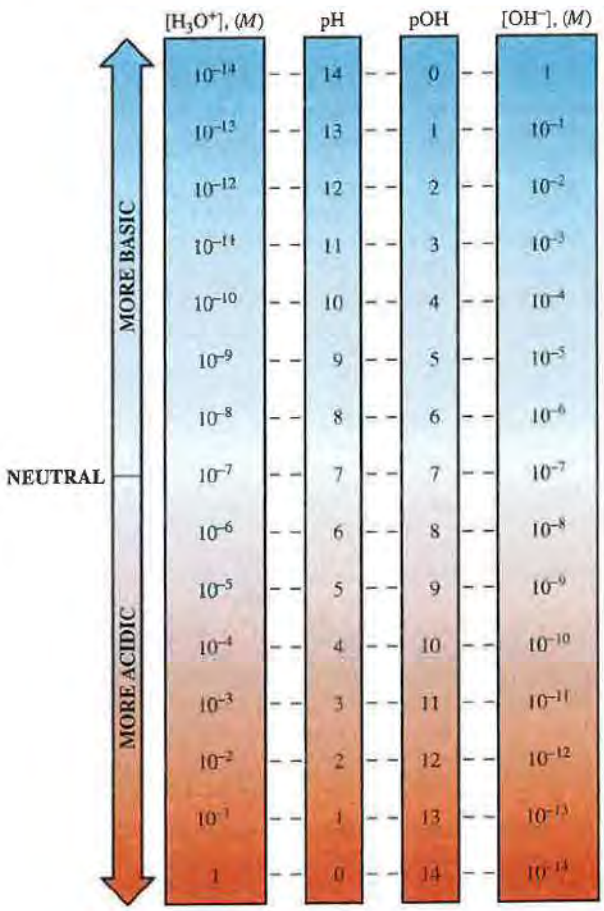


FIGURE 13.16 Notice how these quantities are related. The product of $[H_3O^+]$ and $[OH^-]$ is always 1×10^{-14}, K_w. The sum of the pH and pOH is always 14.



FIGURE 13.17 A pH meter is used to determine the pH of a solution by measuring the voltage that develops when electrodes are dipped into the solution.

Measuring pH

Several methods can be used to measure the pH of a solution. Jake sometimes measures the pH of pool water using a pH meter that is accurate to within hundredths of a pH unit (Figure 13.17). Another less accurate, but often convenient, method is the use of pH *indicators*. Indicators are brightly colored organic dyes that are weak acids or bases. In solution they form an equilibrium with their conjugate bases or conjugate acids, respectively. The color of the indicator depends on whether the dye is in its acidic or basic form.

Consider, for example, the indicator dye phenolphthalein, which is often used in acid-base titrations. Phenolphthalein has an acidic form (abbreviated HIn) that is colorless and a basic form (In^-) that is pink. (The formula HIn is commonly used to represent the acidic form of an indicator. The formula In^- represents the conjugate base of the acid form.)

$$\underset{\text{colorless phenolphthalein}}{HIn(aq)} + H_2O(l) \rightleftharpoons H_3O^+(aq) + \underset{\text{pink phenolphthalein}}{In^-(aq)}$$

Phenolphthalein changes from colorless to pink between pH 8.2 and 10. Below pH 8.2 it is primarily in its colorless acidic form because high concentrations of H_3O^+ cause the equilibrium to lie to the HIn side of the equation. (Recall Le Chatelier's principle from Chapter 12.) Above pH 10 phenolphthalein is primarily in its pink basic form because the low concentration of H_3O^+ causes the equilibrium to lie to the right side of the equation. The color-change ranges for phenolphthalein and other common indicators are shown in Figure 13.18. Note that most of the indicators have acidic and basic forms that are both colored.

An indicator reveals if the pH of a solution is above or below a certain value. Indicators also disclose a specific pH within the indicator's color-change range. For example, the indicator phenol red is sometimes used to find the pH of pool water because its color changes between pH 6.8 and 8.4, the usual pH range for swimming pool water (Figure 13.18). Subtle differences in hues are discernible at slightly different pH values.

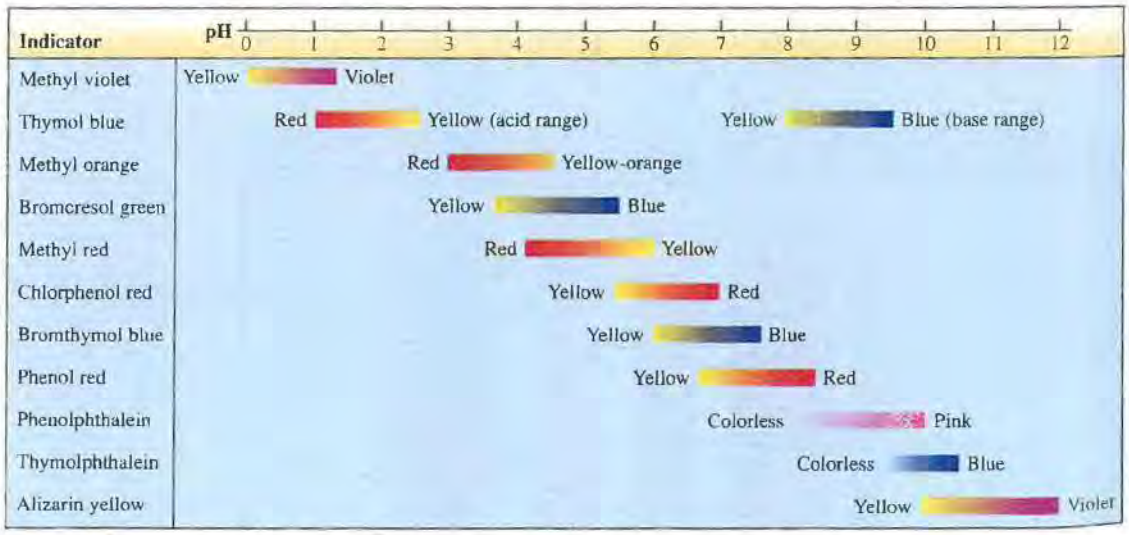


FIGURE 13.18 Each acid-base indicator has a unique pH color-change range. Outside of its range the indicator exists primarily as the acid form or base form, both of which exhibit unique colors. Within its color-change range, the acid and base forms of the indicator are present in similar concentrations. The color of the solution within its range can indicate the pH of the solution.



A



B

FIGURE 13.19 (A) Various types of pH paper contain mixtures of different indicators. They are useful for determining pH at a wide range of pH values. (B) Swimming pool test kits often contain a liquid indicator that can be used to measure the pH.



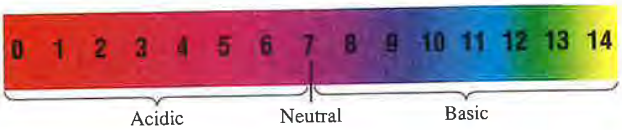


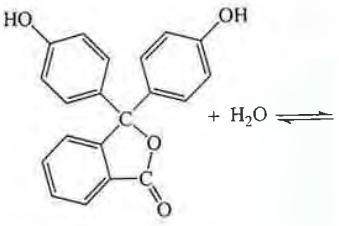
FIGURE 13.20 The juice from a red cabbage can be used as a pH indicator. It exhibits quite a variety of colors under different pH conditions.

Indicators are also useful in measuring pH changes during titration reactions between acids and bases (Chapter 11). For example, consider the titration of an acid with base using phenolphthalein as the indicator. At low pH the phenolphthalein is colorless. As base is added, it reacts with the acid to produce water and the solution becomes less acidic. When the solution just becomes pink, the acid has completely reacted and there is an excess of base, as indicated by the pink color. What would you expect to observe if you ran a titration of a base by adding acid and using methyl red as an indicator?

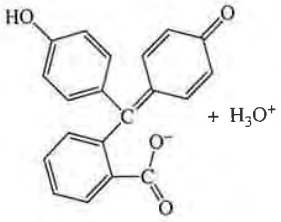
A mixture of indicators having a variety of colors and color-change ranges can be used to measure the pH of any solution. For example, broad-range pH paper is treated with several indicators. The user reads the pH by comparing the color the paper turns to a chart of reference colors and pH values printed on the container (Figure 13.19A). The liquid universal indicator sometimes used in laboratories contains a mixture of indicators and works on the same principle. Liquid indicators are also used in swimming pool test kits (Figure 13.19B).

Another liquid universal indicator can be made from the juice of a red (purple) cabbage. A compound present in the juice is unusual because it can exhibit several different colors based on the pH of the solution being tested (Figure 13.20).

Phenolphthalein, like most acid-base indicators, consists of large organic molecules. The acid form of phenolphthalein is colorless, and the base form is bright pink. Compare the structures of the acid and base forms of phenolphthalein. Can you see that they are a conjugate acid-base pair? Do they differ by one H^+?



Acid, HIn (colorless)



Base, In^- (pink)

13.6 | BUFFERED SOLUTIONS

Olivia's job requires her to find out when there is too much acid or base in garden soil. Because she is considering a career in medicine, she knows what happens when there is too much acid or base in the human bloodstream. Slight changes in blood pH can drastically affect the way our bodies function. For biological reactions to occur normally, blood must stay within the narrow range of pH 7.35 to 7.45. If it rises too high or falls too low, severe health problems, or even death, can occur.

Fortunately our bodies have mechanisms that normally maintain a constant blood pH. One is the action of acid-base buffers in the blood. Blood is a *buffered solution* because it resists changes in pH when limited amounts of acid or base are added to it.

The desired pH of the water in fish tanks depends on the species of fish that live there. Many fish are most healthy in water with a pH of 7.0. Bromthymol blue indicator is often used in test kits for fish tanks. What other indicators shown in Figure 13.18 might help determine if the pH is 7.0, slightly above 7.0, or slightly below 7.0?

ANIMATION: Buffer Solutions

A **buffer**, or a *buffer system*, is a combination of a weak acid and its conjugate base (or a weak base and its conjugate acid) in about equal concentrations.

The buffers in our bloodstream are effective for only normal fluctuations in acid and base concentrations. If too much acid enters the bloodstream, the pH falls below 7.35 and a condition called *acidosis* occurs. *Alkalosis* occurs when the blood is too basic with a pH above 7.45. Both of these situations are dangerous because they disrupt the body's cellular functions.

The *bicarbonate buffer system* regulates pH in the human bloodstream. You might recognize the bicarbonate ion, HCO_3^-, as the conjugate base of carbonic acid, H_2CO_3. Carbonic acid is formed from the reaction of CO_2 with water:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

In aqueous solution, carbonic acid is in equilibrium with its conjugate base:

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

Carbonic acid is also in equilibrium with its conjugate base in the blood. However, the concentration of the bicarbonate ion in the blood is much greater than the concentration that would result from ionization of the acid alone. In the blood, the position of the equilibrium between carbonic acid and the carbonate ion is closely tied to respiration. Changes in respiration can affect blood pH by altering the concentration of CO_2 dissolved in the blood. Consider the combination and simplification of the two reactions just described:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$$

Following Le Chatelier's principle, which was described in Chapter 12, notice how the concentration of CO_2 might affect the blood pH. We would expect an increase in CO_2 concentration to shift the equilibria to the right, causing the $H^+(aq)$ concentration to increase. The result would be a decrease in pH. This is a condition called *respiratory acidosis*, and results from the lungs' inability to release CO_2 at a rate comparable to the rate of its production in the body. This condition occurs in people who suffer from emphysema and other lung conditions.

When the pH of a fish tank is too high or too low, a solution containing an appropriate buffer system is used to bring the pH to its desired level. The $H_2PO_4^-/HPO_4^{2-}$ buffer system is often found in aquarium products that help to maintain a pH 7 environment.

Conversely, a decrease in CO_2 concentration would cause the equilibria to shift to the left, resulting in a higher blood pH. This condition is called *respiratory alkalosis* and occurs when people have panic attacks. During these attacks, they breathe very quickly, releasing more CO_2 than normal, potentially causing the pH of their blood to rise. If these episodes go on for too long, the body's natural reaction will be to cause fainting, which naturally causes breathing to slow down. Have you ever wondered why a remedy for stopping a panic attack is to breathe into a paper bag? Rebreathing the air in the bag increases the concentration of inhaled CO_2, replenishing some that has been lost.

The body's ability to control dramatic changes in pH is based on the naturally occurring amounts of carbonic acid (dissolved CO_2) and bicarbonate ion. This combination of relatively high concentrations of both species keeps blood pH in the range of 7.35 to 7.45.

In the laboratory, chemists often prepare buffered solutions if it is important that they maintain a constant pH in a procedure or experiment. To prepare a solution containing the H_2CO_3/HCO_3^- buffer system, for example, both the acid and the base must be added to the solution in similar amounts. In this case the conjugate base is an ion, so it must be added in the form of a soluble salt such as sodium bicarbonate, $NaHCO_3$. Once dissolved in water, the salt dissociates to $Na^+(aq)$ and $HCO_3^-(aq)$:

$$NaHCO_3(s) \xrightarrow{H_2O} Na^+(aq) + HCO_3^-(aq)$$

ANIMATION: Using a Buffer

How does a buffer system prevent large changes in pH when an acid or base is added to the solution? It does so by reacting with most of the H_3O^+ or OH^- from the added acid or base (Figure 13.21). Consider again the H_2CO_3 HCO_3^- buffer system:

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

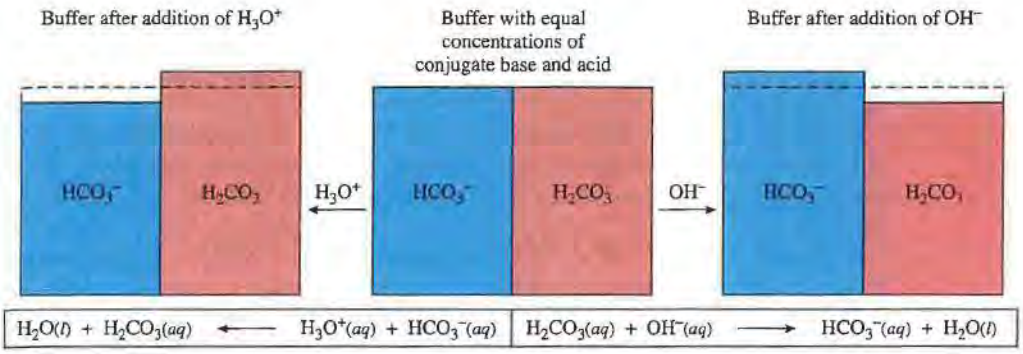


FIGURE 13.21 When an acid is added to an H_2CO_3/HCO_3^- buffer system, most of the acid is consumed by HCO_3^-. When a base is added, most is consumed by H_2CO_3. The relative amounts of the buffer components change, but the pH change is very small.

When a base such as NaOH is added, most of the $OH^-(aq)$ from NaOH reacts with the acid, H_2CO_3:

$$H_2CO_3(aq) + OH^-(aq) \longrightarrow HCO_3^-(aq) + H_2O(l)$$

When an acid such as HCl is added, the $H_3O^+(aq)$ ions from HCl react with the base, HCO_3^-:

$$HCO_3^-(aq) + H_3O^+(aq) \longrightarrow H_2CO_3(aq) + H_2O(l)$$

A buffer system works if there is enough acid and conjugate base in the system to react with all the added $H_3O^+(aq)$ or $OH^-(aq)$. The greater the concentrations of the acid and its conjugate base in the buffer mixture, the more effectively the buffer prevents large changes in pH.

EXAMPLE 13.12 | How a Buffer System Works

Another buffer system in our blood is the $H_2PO_4^-/HPO_4^{2-}$ system. Describe, using a balanced equation, how this system prevents the pH from falling too low when acid enters the bloodstream.

Solution:

The base component of this buffer system, HPO_4^{2-}, reacts with excess H_3O^+ from added acid:

$$HPO_4^{2-}(aq) + H_3O^+(aq) \rightleftharpoons H_2PO_4^-(aq) + H_2O(l)$$

Because most of the excess H_3O^+ is consumed, the pH remains relatively constant.

Practice Problem 13.12

Describe, using a balanced equation, how the $H_2PO_4^-/HPO_4^{2-}$ system in our blood prevents pH levels from rising too high when base enters the bloodstream.

Further Practice: Questions 13.109 and 13.110 at the end of the chapter

Just as a buffer system can moderate pH, the relative amounts of acid and base of a buffer system determine the pH of the solution. In most swimming pool water, for example, an $HOCl/OCl^-$ buffer system is present:

$$HOCl(aq) + H_2O(l) \rightleftharpoons OCl^-(aq) + H_3O^+(aq)$$

Pool water looks best at a pH of 7.5, when the relative amounts of HOCl and OCl^- are equal. At lower pHs, much of the OCl^- is converted to HOCl; at high pHs, much of the HOCl is converted to OCl^-. When the pH is adjusted by adding another kind of acid or base, the equilibrium shifts, so the relative amounts of HOCl and OCl^- also adjust.

The examples of buffer systems we have discussed are conjugate pairs of weak acids and weak bases. Why wouldn't a strong acid, like HCl, and its conjugate base make an effective buffer? Strong acids ionize completely in water:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

The Cl^- ion does not react with H_3O^+ to form HCl, because HCl is a strong acid and the acid molecule does not exist in solution. For a buffer system to work, both of the components of the conjugate acid-base pair must exist in equilibrium in solution. This is not the case with a strong acid or base.

EXAMPLE 13.13 | Buffer Systems

Which of the following, when added to water, can act as a buffer system? For each buffer system, write a balanced equation.

(a) HF and NaF
(b) HNO_3 and KNO_3
(c) NH_3 and NH_4Cl

Solution:

(a) This is a buffer system, because HF is a weak acid and the ionic compound NaF provides a source of its conjugate base, F^-, when it dissolves in water. We can ignore the spectator ion, Na^+. The following equation shows the conjugate acid-base pair that exists in solution at equilibrium:

$$HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$$

(b) This is not a buffer system, because HNO_3 is a strong acid. Strong acids cannot be components of buffers, because they ionize completely in water and are not in equilibrium with their conjugate bases.

(c) This is a buffer system, because NH_3 is a weak base and its conjugate acid, NH_4^+, forms when the ionic compound NH_4Cl dissolves in water. We can ignore the spectator ion, Cl^-. The following is a conjugate acid-base pair that will exist at equilibrium in solution:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Practice Problem 13.13

Which of the following systems, when added to water, can act as a buffer system? For each buffer system, write a balanced equation.

(a) HCl and NaOH
(b) CH_3CO_2H and $NaCH_3CO_2$
(c) HBr and KBr

Further Practice: Questions 13.111 and 13.112 at the end of the chapter

SUMMARY

Acids and bases are important in our everyday lives. They contribute to food tastes, have properties that are helpful for cleaning, and participate in many chemical processes. The Brønsted-Lowry theory is most widely used to describe the behavior of acids and bases. According to this theory, an acid is a substance that donates an H^+ ion to another substance; a base is an H^+ acceptor. Acids and bases can be ranked according to their strength, or how readily they donate or accept H^+ ions. When strong acids dissolve in water, they completely ionize, so essentially no acid molecules remain unreacted. When a weak acid dissolves in water, only a small percentage of the acid molecules ionize. Most remain as molecules. Equilibrium constants called acid ionization constants, K_a, can be used to compare the strengths of different weak acids. The stronger the acid, the larger the K_a value.

The relative concentrations of H_3O^+ and OH^- ions in an aqueous solution at 25°C are determined by the ion-product constant for water, K_w:

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \quad K_w = 1.0 \times 10^{-14}$$
$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

Neutral solutions have equal concentrations of H_3O^+ and OH^-. Acidic solutions contain a greater concentration of H_3O^+ than OH^-. Basic solutions contain a greater concentration of OH^- than H_3O^+. The acidity of a solution is commonly expressed in terms of pH, the negative logarithm of the H_3O^+ concentration:

$$pH = -\log[H_3O^+]$$

At 25°C, acidic solutions have a pH less than 7, basic solutions have a pH greater than 7, and neutral solutions have a pH equal to 7.

Buffered solutions contain a weak acid and its conjugate base (or a weak base and its conjugate acid) in similar concentrations. Buffers help to prevent large changes in pH by reacting with small amounts of added acid or base.

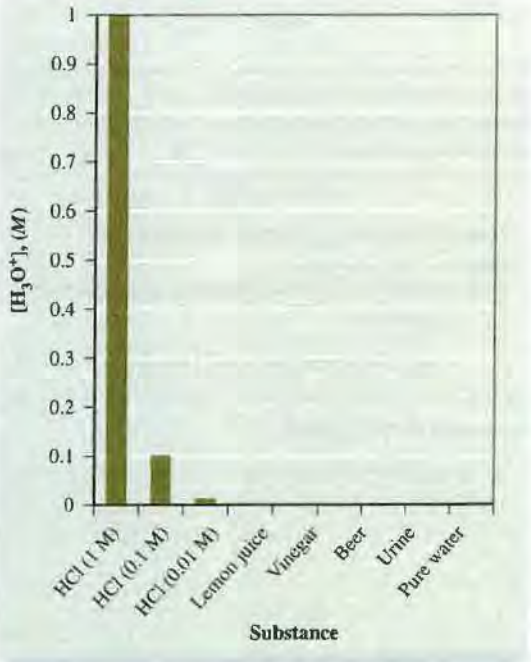
Math Toolbox 13.1 | Log and Inverse Log Functions

Log Functions

Let's consider several substances that vary in their $H_3O^+(aq)$ concentrations:

Substance	$[H_3O^+]$
HCl (1 *M*)	1 *M*
HCl (0.1 *M*)	0.1 *M* (or 10^{-1} *M*)
HCl (0.01 *M*)	0.01 *M* (or 10^{-2} *M*)
lemon juice	0.001 *M* (or 10^{-3} *M*)
vinegar	0.0001 *M* (or 10^{-4} *M*)
beer	0.00001 *M* (or 10^{-5} *M*)
urine	0.000001 *M* (or 10^{-6} *M*)
pure water	0.0000001 *M* (or 10^{-7} *M*)

Except for pure water, these substances are all acids. Suppose we wanted to compare their concentrations graphically. In a bar graph, notice what happens to those with the lower concentrations:



Math Toolbox 13.1 (*continued*)

Can you distinguish lemon juice, which we can consume, from stomach acid, which causes damage to the esophagus? No. Logarithms are very useful in situations like this because logarithms make it easier to compare numbers that are inconvenient to work with and differ by orders of magnitude.

The *log* function on your calculator stands for *logarithm to the base 10*. When you take the log of a number, you are asking, "10 to what power will give the value after the word *log*?" For example, the log of 1000 is 3, because 10 must be raised to the third power to get 1000:

$$\log(1000) = 3.0$$

Likewise, the log of 0.001 is −3 because 10 must be raised to the −3 power to get 0.001:

$$\log(0.001) = -3.0$$

To use the log function on most calculators, you first press the log key and then enter the number. (On some calculators, these two steps may be reversed.) Try taking the log of 0.0010 on your calculator. You should get −3.0.

Step 1: Press the log key.
Step 2: Enter 0.001 (or 1×10^{-3}), and then the ENTER or = key.

You should get −3.0. If this does not work, try reordering the key strokes.

Numerical values with large positive or negative exponents, such as concentrations of H_3O^+, are inconvenient to work with. A more convenient way to express such values is to use a logarithmic scale. The pH scale is an example of a logarithmic scale. By convention, the "p" of something is its negative logarithm:

$$pX = -\log X$$

In chemistry, the most common examples are the pH scale and the pOH scale:

$$pH = -\log[H^+] = -\log[H_3O^+]$$
$$pOH = -\log[OH^-]$$

To obtain the pH or pOH of a solution, you must first take the log of the concentration and then multiply by −1 (change the sign). Try using your calculator to determine the pH of a solution that has an H_3O^+ concentration of 5.0×10^{-4} *M*.

$$pH = -\log[H_3O^+]$$

Step 1: Press the +/− (change of sign) key.
Step 2: Press the log key.
Step 3: Enter the H_3O^+ concentration, and then the ENTER or = key.

$$pH = -\log(5.0 \times 10^{-4}) = 3.30$$

If you did not get the right answer, your calculator may require you to reorder the key strokes.

EXAMPLE 13.14 Log Functions

Use your calculator to find the pH of the following solutions.

(a) $[H_3O^+] = 1.0 \times 10^{-8}$ *M*
(b) $[H_3O^+] = 6.2 \times 10^{-1}$ *M*

Solution:

To calculate the pH from the H_3O^+ concentration, use the following relationship:

$$pH = -\log[H_3O^+]$$

Using your calculator, first enter a negative sign, followed by the log key and the H_3O^+ concentration. Then hit the ENTER or = key.

(a) $pH = -\log(1.0 \times 10^{-8}) = 8.00$
(b) $pH = -\log(6.2 \times 10^{-1}) = 0.21$

Practice Problem 13.14

Use your calculator to find the pH of the following solutions.

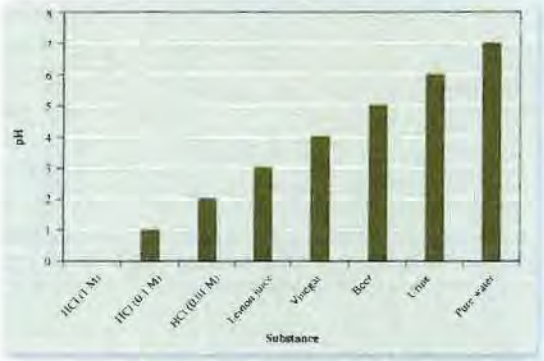
(a) $[H_3O^+] = 7.5 \times 10^{-3}$ *M*
(b) $[H_3O^+] = 5.6 \times 10^{-8}$ *M*

Further Practice: Questions 13.3 and 13.4 at the end of the chapter

If we take the log of our original data and then change the sign, we get the following values:

Substance	$[H_3O^+]$	$-\log[H_3O^+]$
HCl (1 *M*)	1 *M*	0
HCl (0.1 *M*)	0.1 *M* (or 10^{-1} *M*)	1
HCl (0.01 *M*)	0.01 *M* (or 10^{-2} *M*)	2
lemon juice	0.001 *M* (or 10^{-3} *M*)	3
vinegar	0.0001 *M* (or 10^{-4} *M*)	4
beer	0.00001 *M* (or 10^{-5} *M*)	5
urine	0.000001 *M* (or 10^{-6} *M*)	6
pure water	0.0000001 *M* (or 10^{-7} *M*)	7

A plot of the log values now allows us to distinguish the H_3O^+ content in each of these substances:



Math Toolbox 13.1 (*continued*)

The pH scale is more convenient to use. However, we have to keep in mind that an integer difference in pH is a 10-fold difference in H_3O^+ concentration. For example, the pH of beer is around 5 and that of vinegar is around 4. Vinegar is 10 times as acidic as beer (10^{-4} for vinegar versus 10^{-5} for beer).

Inverse Log (10^x) Functions

When you have an equation that involves a log function, you sometimes need to solve for the quantity that follows the log. For example, if you are given pH, you might want to solve for $[H_3O^+]$:

$$pH = -\log[H_3O^+]$$

To solve for $[H_3O^+]$, we must take the inverse log (antilog) of both sides of the equation. To rearrange the equation so that $[H_3O^+]$ is on one side by itself, first multiply both sides by −1:

$$-pH = \log[H_3O^+]$$

To get rid of the log expression in front of $[H_3O^+]$, take the inverse log of both sides:

$$\text{inverse log}\,(-pH) = \text{inverse log}\,(\log[H_3O^+])$$

The value of $[H_3O^+]$, then, equals the inverse log of the negative pH:

$$\text{inverse log}\,(-pH) = [H_3O^+]$$

Taking the inverse log of –pH is the same as 10 raised to the –pH power. The inverse log function on your calculator is the 10^x function, which is often found above the log function as an alternate function for the same key. Another way to describe the relationship between $[H_3O^+]$ and pH is to say $[H_3O^+]$ equals 10 raised to the –pH power:

$$[H_3O^+] = 10^{-pH}$$

To use the inverse log (10^x) function for this calculation on most calculators, you first press the INV, SHIFT, or 2nd button followed by the log key. Then enter the negative of the pH value. The inverse log function, or 10^x function, should appear on your calculator above the log function as the alternative function. By pressing INV, SHIFT, or 2nd button first, you tell the calculator to use the alternative function. (On some calculators, you have to enter the –pH value first.) Try taking the inverse log where pH is 3.00.

$$[H_3O^+] = 10^{-pH}$$

Step 1: Press the INV, SHIFT, or 2nd button.
Step 2: Press the log button to use the 10^x function.
Step 3: Enter –3.00, and then the ENTER or = key.

Did you get $[H_3O^+] = 0.0010$ or 1.0×10^{-3}? If not, your calculator may require you to reorder the key strokes. The steps for calculating $[OH^-]$ from pOH are the same as calculating $[H_3O^+]$ from pH:

$$[H_3O^+] = 10^{-pH}$$

$$[OH^-] = 10^{-pOH}$$

The general calculation steps for calculating $[H_3O^+]$ or $[OH^-]$ are

Step 1: Press the INV, SHIFT, or 2nd button.
Step 2: Press the log button.
Step 3: Enter –pH (or –pOH), and then the ENTER or = key.

(On some calculators you may have to reorder the key strokes.)

EXAMPLE 13.15 Inverse Log Functions

Calculate the H_3O^+ concentrations for the solutions with the following pH values.

(a) pH = 9.40
(b) pH = 3.60

Solution:

To calculate the H_3O^+ concentration from the pH, use the following relationship:

$$[H_3O^+] = 10^{-pH}$$

Using your calculator, first enter the INV, SHIFT, or 2nd button. Then press the log key to give you the 10^x function. (On some calculators the 10^x button is its own key.) Finally, press the +/– (change sign) key, enter the pH, and then the ENTER or = key.

(a) $[H_3O^+] = 10^{-9.40} = 4.0 \times 10^{-10}$
(b) $[H_3O^+] = 10^{-3.60} = 2.5 \times 10^{-4}$

Practice Problem 13.15

Calculate the OH^- concentration from the following pOH values.

(a) pOH = 8.00
(b) pOH = 2.60

Further Practice: Questions 13.5 and 13.6 at the end of the chapter

KEY RELATIONSHIPS

Relationship	Equation
The self-ionization constant for water equals the product of the hydronium ion concentration and the hydroxide ion concentration. At 25°C, this constant equals 1.0×10^{-14}.	$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$
The pH of a solution equals the negative log of the hydronium ion concentration. The pOH of a solution equals the negative log of the hydroxide ion concentration.	$pH = -\log[H_3O^+]$ $pOH = -\log[OH^-]$
The hydronium ion concentration equals the inverse log of the negative pH value. The hydroxide ion concentration equals the inverse log of the negative pOH value.	$[H_3O^+] = 10^{-pH}$ $[OH^-] = 10^{-pOH}$
The pH and pOH of a solution total 14.00.	$pH + pOH = 14.00$

KEY TERMS

acid ionization constant, K_a (13.3)
acidic solution (13.4)
amphoteric substance (13.1)
Arrhenius model of acids and bases (13.1)
basic solution (13.4)
Brønsted-Lowry theory (13.1)
buffer (13.6)
conjugate acid (13.1)
conjugate base (13.1)
hydronium ion (13.1)
ion-product constant of water, K_w (13.4)
neutral solution (13.4)
polyprotic acid (13.3)
pH (13.5)
self-ionization (13.4)
strong acid (13.2)
strong base (13.2)
weak acid (13.2)
weak base (13.2)

QUESTIONS AND PROBLEMS

The following questions and problems, except for those in the *Additional Questions* section, are paired. Questions in a pair focus on the same concept. Answers to the odd-numbered questions and problems are in Appendix D.

Matching Definitions with Key Terms

13.1 Match the key terms with the descriptions provided.
(a) a base that ionizes or dissociates completely when dissolved in water
(b) a theory that defines an acid as a substance that donates an H^+ to another substance in solution and a base as a substance that accepts an H^+ in solution
(c) a substance that forms after a base gains an H^+ ion
(d) an acid containing more than one acidic hydrogen
(e) a solution in which the H_3O^+ concentration is greater than the OH^- concentration; a solution with a pH less than 7
(f) a base that does not completely ionize when dissolved in water
(g) a substance that can act as either an acid or a base
(h) the equilibrium constant for the self-ionization of water; the product of the H_3O^+ ion concentration and the OH^- ion concentration in any aqueous solution
(i) a process in which one molecule transfers a proton to another molecule of the same substance; water does this to a very small extent

13.2 Match the key terms with the descriptions provided.
(a) an acid that ionizes completely when dissolved in water
(b) a model that describes acids as substances that generate H^+ ions in solution and bases as substances that generate OH^- ions in solution
(c) an aqueous hydrogen ion; $H_3O^+(aq)$
(d) a solution in which the OH^- concentration is greater than the H_3O^+ concentration; a solution with a pH greater than 7
(e) an acid that does not completely ionize when dissolved in water
(f) a substance that forms after an acid loses an H^+ ion
(g) an equilibrium constant for the ionization of an acid in water; a value that expresses the strength of a weak acid
(h) a measure of the acidity of aqueous solutions; the negative logarithm of the H_3O^+ concentration
(i) a combination of a weak acid and its conjugate base (or a weak base and its conjugate acid) in similar amounts; when in solution, something that helps prevent large changes in pH when small amounts of $H_3O^+(aq)$ or $OH^-(aq)$ are added

Math Toolbox Questions

13.3 Use your calculator to find the log of the following numbers.
(a) 10^{-9}
(b) 1×10^{-11}
(c) 7.4×10^{3}
(d) 10^{5}
(e) 1

13.4 Use your calculator to find the log of the following numbers.
(a) 10^{4}
(b) 1×10^{-6}
(c) 1.7×10^{8}
(d) 10^{-8}
(e) 10

13.5 Use your calculator to find the inverse log of the following numbers.
(a) 1.2
(b) –6.2
(c) 0

13.6 Use your calculator to find the inverse log of the following numbers.
(a) 12.7
(b) –9.4
(c) 1

What Are Acids and Bases?

13.7 What are some properties of acids?

13.8 What are some properties of bases?

13.9 List some common foods or household products that are bases.

13.10 List some common foods or household products that are acids.

13.11 In terms of the Arrhenius concept of acids and bases, how does an acid behave when dissolved in water? How does an Arrhenius base behave?

13.12 How does the Arrhenius concept of acids and bases emphasize that they are electrolytes?

13.13 How is the Brønsted-Lowry theory of acids and bases different from the Arrhenius model?

13.14 Why is $H_3O^+(aq)$ a more accurate representation of an aqueous H^+ ion than $H^+(aq)$?

13.15 Identify the first reactant in each equation as an acid or a base.
(a) $HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$
(b) $SO_4^{2-}(aq) + H_2O(l) \rightleftharpoons HSO_4^-(aq) + OH^-(aq)$
(c) $C_6H_5OH(aq) + NaOH(aq) \rightleftharpoons H_2O(l) + C_6H_5O^-(aq) + Na^+(aq)$

13.16 Identify the first reactant in each equation as an acid or a base.
(a) $CH_3CH_2NH_3^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CH_2NH_2(aq)$
(b) $HCN(aq) + H_2O(l) \rightleftharpoons CN^-(aq) + H_3O^+(aq)$
(c) $NaOH(aq) + CH_3CO_2H(aq) \rightleftharpoons H_2O(aq) + CH_3CO_2^-(aq) + Na^+(aq)$

13.17 Write the formula of the conjugate base of each acid.
(a) HNO_2
(b) HF
(c) H_3BO_3

13.18 Write the formula of the conjugate base of each acid.
(a) HOCl
(b) H_3PO_4
(c) HCO_3^-

13.19 Write the formula of the conjugate acid of each base.
(a) OH^-
(b) $C_6H_5NH_2$
(c) HCO_3^-

13.20 Write the formula of the conjugate acid of each base.
(a) NH_3
(b) F^-
(c) CO_3^{2-}

13.21 Write an equation to show how the following acids or bases react with water.
(a) $HCl(g)$
(b) $HClO_4(l)$
(c) $CH_3CO_2^-(aq)$

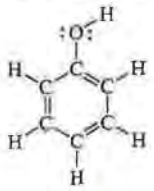
13.22 Write an equation to show how the following acids or bases behave in water.
(a) $NaOH(s)$
(b) $CN^-(aq)$
(c) $HNO_3(l)$

13.23 Which of the following are amphoteric species?
(a) H_2O
(b) HSO_3^-
(c) SO_4^{2-}

13.24 Which of the following are amphoteric species?
(a) NH_4^+
(b) HBO_3^{2-}
(c) $H_2PO_4^-$

13.25 In an H_2CO_3 molecule, both hydrogen atoms are bonded to oxygen atoms. How many acidic hydrogen atoms does carbonic acid have?

13.26 The Lewis structure for the acid phenol is shown. How many acidic hydrogen atoms does phenol have?



Strong and Weak Acids and Bases

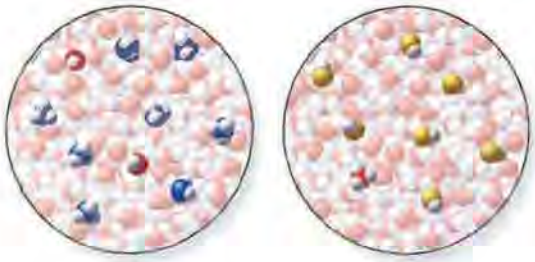
13.27 How do strong acids and bases differ from weak acids and bases?

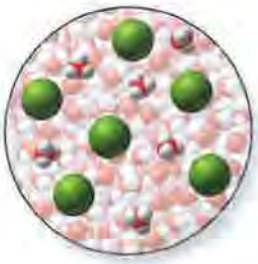
13.28 What do we mean when we say weak acids and bases only partially ionize in water?

13.29 Identify each of the following as a strong acid, weak acid, strong base, or weak base. For each, write an equation that describes its reaction with water.
(a) $H_2SO_4(aq)$
(b) $Ca(OH)_2(aq)$
(c) $Na_2CO_3(aq)$
(d) $H_3C_6H_5O_7(aq)$
(e) $C_6H_5NH_2(aq)$

13.30 Identify each of the following as a strong acid, weak acid, strong base, or weak base. For each, write an equation that describes its reaction with water.
(a) $HCN(aq)$
(b) $NaCN(aq)$
(c) $LiOH(aq)$
(d) $CH_3NH_2(aq)$
(e) $H_2C_2O_4(aq)$

13.31 Match the molecular-level diagram to each of the following compounds in aqueous solution: HCl, HF, and NH_3.





13.32 Draw a molecular-level representation of a solution of each of the following acids and bases.
(a) $Ba(OH)_2$
(b) NaOH
(c) HOCl

13.33 Which compound, KBr or KF, is expected to behave as a weak base? Why?

13.34 Which compound, $NaHCO_3$ or $NaNO_3$, is expected to behave as a weak base? Why?

Relative Strengths of Weak Acids

13.35 If acid A ionizes to a greater extent than acid B, which acid is stronger?

13.36 If acid C has a smaller K_a value than acid D, which acid is stronger?

13.37 Which solution has the highest concentration of H_3O^+, a 0.50 M solution of acetic acid, CH_3CO_2H, or a 0.50 M solution of formic acid, HCO_2H? Consult Table 13.5 for K_a values.

13.38 Which acid ionizes to the greater extent in water, nitrous acid, HNO_2, or hydrofluoric acid, HF? Consult Table 13.5 for K_a values.

13.39 Sodium fluoride, NaF, and sodium acetate, $NaCH_3CO_2$, are ionic compounds that contain the conjugate bases of weak acids. Assuming you start with solutions of equal concentrations, which would have the greater concentration of OH^-? (*Hint:* Consult Table 13.5.)

13.40 Potassium hypochlorite, KOCl, and potassium nitrite, KNO_2, are ionic compounds that contain the conjugate bases of weak acids. Assuming you start with solutions of equal concentrations, which would have the greater concentration of OH^-? (*Hint:* Consult Table 13.5.)

13.41 What is the difference between a concentrated acid solution and a strong acid solution?

13.42 Can a weak acid solution have a higher H_3O^+ concentration than a strong acid solution? Explain.

13.43 What is a polyprotic acid?

13.44 Write the formulas and names of three different polyprotic acids.

13.45 Why does hydrosulfuric acid, H_2S, have two acid ionization constants? Write equations that correspond to each K_a value.

13.46 Which species, H_2S or HS^-, ionizes to the greater extent in water? Explain.

13.47 Which molecules or ions are present in an aqueous solution of oxalic acid, $H_2C_2O_4$, a polyprotic acid? Which is present in the least concentration?

13.48 Which molecules or ions are present in an aqueous solution of citric acid, $H_3C_6H_5O_7$, a triprotic acid? Which is present in the smallest concentration?

Acidic, Basic, and Neutral Solutions

13.49 Under what conditions is the ion-product constant of water, K_w, a constant?

13.50 Does the product of the H_3O^+ ion concentration and the OH^- ion concentration equal 1.0×10^{-14} when substances such as acids and bases are added to water? Explain.

13.51 What are the concentrations of H_3O^+ and OH^- in pure water (at 25°C)?

13.52 What happens to the concentrations of H_3O^+ and OH^- when acid is added to water? When a base is added to water?

13.53 Identify each of the following conditions as acidic, basic, or neutral.
(a) $[H_3O^+] < [OH^-]$
(b) $[H_3O^+] = 1.0 \times 10^{-4}\ M$
(c) $[OH^-] = 1.0 \times 10^{-3}\ M$

13.54 Identify each of the following conditions as acidic, basic, or neutral.
(a) $[OH^-] = 1.0 \times 10^{-11}\ M$
(b) $[H_3O^+] = 1.0 \times 10^{-9}\ M$
(c) $[H_3O^+] > [OH^-]$

13.55 Calculate the H_3O^+ ion concentration from the OH^- ion concentration. Then identify the solution as acidic, basic, or neutral.
(a) $[OH^-] = 1.0 \times 10^{-3}\ M$
(b) $[OH^-] = 1.0 \times 10^{-11}\ M$
(c) $[OH^-] = 3.2 \times 10^{-8}\ M$

13.56 Calculate the OH^- ion concentration from the H_3O^+ ion concentration. Then identify the solution as acidic, basic, or neutral.
(a) $[H_3O^+] = 1.0 \times 10^{-3}\ M$
(b) $[H_3O^+] = 1.0 \times 10^{-5}\ M$
(c) $[H_3O^+] = 7.3 \times 10^{-10}\ M$

13.57 What is the concentration of H_3O^+ in each of the following solutions?
(a) 0.010 *M* HNO_3
(b) 0.020 *M* $HClO_4$
(c) 0.015 *M* NaOH

13.58 What is the concentration of OH^- in each of the following solutions?
(a) 0.0010 *M* KOH
(b) 0.0050 *M* NaOH
(c) 0.0010 *M* HCl

The pH Scale

13.59 Olivia noticed that the pH of her garden soil decreased after it rained. Did her garden soil become more acidic or less acidic?

13.60 Jake noticed that swimming pool water tends to increase in pH more quickly during hot weather. Does pool water tend to become more acidic or more basic during hot weather?

13.61 If one solution has an H_3O^+ concentration that is 10 times that of another solution, what is the difference in their pH values?

13.62 If solution A has a pH that is three pH units greater than that of solution B, how much greater is the H_3O^+ concentration in solution B than in solution A?

13.63 Why does the pH scale typically range from pH 0 to pH 14?

13.64 Can a solution have a pH less than 0 or greater than 14? Explain.

13.65 Can the pH of an acid solution be greater than 7?

13.66 What is the pH range for acidic solutions? For basic solutions?

13.67 What is the pH of solutions having the following H_3O^+ concentrations? Identify each as acidic, basic, or neutral.
(a) $[H_3O^+] = 1.0 \times 10^{-3}\ M$
(b) $[H_3O^+] = 1.0 \times 10^{-13}\ M$
(c) $[H_3O^+] = 3.4 \times 10^{-10}\ M$

13.68 What is the pH of solutions having the following H_3O^+ concentrations? Identify each as acidic, basic, or neutral.
(a) $[H_3O^+] = 1.0 \times 10^{-9}\ M$
(b) $[H_3O^+] = 1.0 \times 10^{-14}\ M$
(c) $[H_3O^+] = 7.2 \times 10^{-3}\ M$

13.69 What is the pH of solutions having the following OH^- concentrations? Identify each as acidic, basic, or neutral.
(a) $[OH^-] = 1.0 \times 10^{-4}\ M$
(b) $[OH^-] = 1.0 \times 10^{-7}\ M$
(c) $[OH^-] = 8.2 \times 10^{-10}\ M$

13.70 What is the pH of solutions having the following OH^- concentrations? Identify each as acidic, basic, or neutral.
(a) $[OH^-] = 6.4 \times 10^{-4}\ M$
(b) $[OH^-] = 1.0\ M$
(c) $[OH^-] = 2.7 \times 10^{-10}\ M$

13.71 What is the pH of the following solutions? Identify each as acidic, basic, or neutral.
(a) 0.010 *M* HNO_3
(b) 0.020 *M* $HClO_4$
(c) 0.015 *M* NaOH

13.72 What is the pH of the following solutions? Identify each as acidic, basic, or neutral.
(a) 0.0010 *M* KOH
(b) 0.0050 *M* NaOH
(c) 0.0010 *M* HCl

13.73 Complete the following table.

	$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic or Basic?
(a)	1.0×10^{-5}				
(b)		1.0×10^{-4}			
(c)				8.00	
(d)			8.54		
(e)		9.0×10^{-10}			

13.74 Complete the following table.

	$[H_3O^+]$	$[OH^-]$	pH	pOH	Acidic or Basic?
(a)			0.40		
(b)		9.0×10^{-10}			
(c)	1.0×10^{-8}				
(d)				2.00	
(e)	4.5×10^{-2}				

13.75 Determine the pH and pOH of a 0.0050 *M* HCl solution. What is the relationship between the pH and pOH values?

13.76 Determine the pH and pOH of a 0.050 *M* NaOH solution. What is the relationship between the pH and pOH values?

13.77 Determine the H_3O^+ concentration in solutions with each of the following pH values. Identify each solution as acidic, basic, or neutral.
(a) pH = 5.00
(b) pH = 12.00
(c) pH = 5.90

13.78 Determine the OH^- concentration given the pH of the following solutions. Identify each solution as acidic, basic, or neutral.
(a) pH = 5.00
(b) pH = 11.00
(c) pH = 12.80

13.79 Determine the H_3O^+ concentration in the following solutions. Identify each solution as acidic, basic, or neutral.
(a) household ammonia, pH = 11.00
(b) blood, pH = 7.40
(c) lime juice, pH = 1.90

13.80 Determine the H_3O^+ concentration in the following solutions. Identify each solution as acidic, basic, or neutral.
(a) orange juice, pH = 3.50
(b) lye, 4.0% sodium hydroxide, pH = 14.00
(c) saliva, pH = 7.00

13.81 What is the OH^- concentration in each of the following solutions?
(a) gastric juice, pH = 1.00
(b) Milk of Magnesia, pH = 10.50
(c) Seven-Up, pH = 3.60

13.82 What is the H_3O^+ concentration in each of the following solutions?
(a) lake water, pOH = 6.00
(b) coffee, pOH = 8.90
(c) borax, pOH = 4.50

13.83 Would you expect the pH of 0.010 *M* acetic acid to be higher or lower than 2.0?

13.84 Would you expect the pH of a 0.010 *M* NH_3 solution to be higher or lower than 12.0?

13.85 If the pH of an NaOH solution is 13.0, what is the concentration of NaOH?

13.86 If the pH of an HCl solution is 2.0, what is the concentration of HCl?

13.87 If you need to know the pH of a solution to two decimal places, which method of pH measurement discussed in this chapter would be the best choice?

13.88 What are the advantages of using pH paper to measure pH?

13.89 What types of compounds are acid-base indicators?

13.90 Why do we use the symbols HIn and In^- to represent the different forms of an indicator? Which is the acid form and which is the base form of the indicator?

13.91 What is a pH color-change range for an indicator? How is it useful for determining pH between the range of pH values?

13.92 How can an indicator have two color-change ranges? What characteristic must be present?

13.93 Which of the indicators listed in Figure 13.18 would be a good choice for determining the pH of a solution that normally has a pH between 9 and 10?

13.94 Which of the indicators in Figure 13.18 would be a good choice for determining the pH of a solution that normally has a pH between 4 and 5?

13.95 Predict the color of the methyl orange indicator in a solution that has a pH of 3.5.

13.96 Predict the color of the thymol blue indicator in a solution that has a pH of 8.8.

13.97 What can we say about the pH of a solution that is red when the indicator thymol blue is added? When the solution is yellow? When the solution is blue?

13.98 What can we say about the pH of a solution that is colorless when phenolphthalein is added? When the solution is dark pink?

13.99 What is a universal indicator? How are universal indicator and pH paper similar?

13.100 How is litmus paper different from pH paper?

13.101 What would you expect to observe if you ran a titration of an acid at a pH of 1.0 by adding base and using bromocresol green as an indicator?

13.102 What would you expect to observe if you ran a titration of a base at a pH of 11.0 by adding acid and using methyl red as an indicator?

Buffered Solutions

13.103 How is a buffered solution different from a weak acid or a weak base dissolved in water?

13.104 What types of substances make good buffers?

13.105 What is the role of a buffer system in human blood?

13.106 How does the pH change when an acid or base is added to a buffered solution?

13.107 If CH_3CO_2H is added to water, identify an ionic compound that could be added to make a buffered solution.

13.108 If Na_2SO_4 is added to water, what should also be added to make a buffered solution?

13.109 A buffer system used to maintain solution pH values around 12 to 13 is prepared by adding similar concentrations of Na_2HPO_4 and Na_3PO_4 to water.
(a) Write a balanced equation showing the acid and conjugate base in equilibrium. Omit spectator ions.
(b) Describe how this buffer system prevents large pH changes when an acid is added.

13.110 A buffer system used to maintain solution pH values around 5 is prepared by adding similar concentrations of CH_3CO_2H and $NaCH_3CO_2$ to water.
(a) Write a balanced equation showing the acid and conjugate base in equilibrium. Omit spectator ions.
(b) Describe how this buffer system prevents large pH changes when a base is added.

13.111 Which of the following, when added to water, form a buffered solution?
(a) HOCl and NaCl
(b) HNO_2 and KNO_2
(c) CH_3NH_2 and CH_3NH_3Cl

13.112 Which of the following, when added to water, form a buffered solution?
(a) HF and NaF
(b) HI and KI
(c) HOCl and NaOCl

13.113 The bicarbonate buffer system operates in the body's extracellular fluid by the following process:

$$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H^+(aq)$$

What is expected to happen to the blood pH if the lungs are unable to expel CO_2 at the same rate as the body produces it? Describe what happens in terms of Le Chatelier's principle.

13.114 For the bicarbonate buffer system described in Question 13.113, what is expected to happen to the blood pH during a panic attack in which more CO_2 than normal is released by the body? Describe what happens in terms of Le Chatelier's principle.

Additional Questions

13.115 Why is NH_3 a base but CH_4 is not? Is it possible for CH_4 to act as a Brønsted-Lowry base?

13.116 Are both acid concentration and acid strength important in determining the concentration of H_3O^+ in an acid solution? Explain.

13.117 Describe what happens when NaF dissolves in water. Describe the acid or base properties of this solution.

13.118 Write an equation that shows what happens when $Ba(OH)_2$ dissolves in water. What is the pH of 0.01 *M* $Ba(OH)_2$?

13.119 Can the pH of a solution of NaOH ever be less than 7 at 25°C? Explain.

13.120 Can the pH of a solution of HCl ever be greater than 7 at 25°C? Explain.

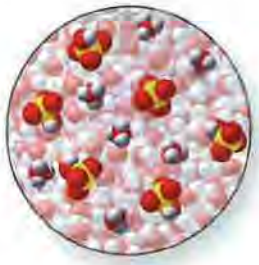
13.121 Which should have the lower pH, a 0.10 *M* H_2SO_4 solution or a 0.10 *M* HNO_3 solution? Explain.

13.122 Consider the reaction classifications we studied in Chapter 5. For each reaction type given, write a chemical reaction in which HCl is a reactant.
(a) neutralization (double-displacement)
(b) single-displacement

13.123 Water in a properly maintained swimming pool with a pH of 7.5 has approximately equal concentrations of $HOCl(aq)$ and $OCl^-(aq)$.
(a) Write a balanced equation showing the acid and conjugate base in equilibrium.
(b) Describe what happens to this equilibrium when an acid, such as HCl, is added? What happens to the relative concentrations of HOCl and OCl^-? What happens to the pH of the pool water?

13.124 Explain why ions such as Cl^-, NO_3^-, and I^- do not act as bases in water.

13.125 Label all of the species shown in this image of a sulfuric acid, $H_2SO_4(aq)$, solution. Explain why there are no H_2SO_4 molecules shown.



13.126 Formic acid, HCO_2H, is the substance present in fire ants that causes a sting when they bite you. What is the conjugate base of formic acid?

13.127 Suppose you wanted to prepare a buffer solution with formic acid, HCO_2H, a weak acid with $K_a = 1.8 \times 10^{-4}$. What substances might you mix together to prepare the buffer?

13.128 Ascorbic acid (vitamin C) is a diprotic acid with the formula $H_2C_6H_6O_6$. The equilibrium constants for the two ionization processes associated with ascorbic acid are

$$K_{a1} = 9.1 \times 10^{-4}$$
$$K_{a2} = 2.0 \times 10^{-12}$$

Write the equilibrium equations that correspond to the ionization processes for ascorbic acid. Rank the species present in a solution of ascorbic acid from lowest to highest concentration.

CHAPTER

16

Organic Chemistry

16.1 Representations of Organic Molecules
16.2 Hydrocarbons
16.3 Acyclic Hydrocarbons
16.4 Cyclic Hydrocarbons
16.5 Alcohols and Ethers
16.6 Aldehydes and Ketones
16.7 Carboxylic Acids and Esters
16.8 Amines
16.9 Organic Nomenclature
Summary
Key Terms
Questions and Problems



When Scott and Jennifer arrive early for their chemistry class, they find an assignment written on the board: "Identify a molecule containing at least eight atoms that is important to the career you are working toward. Describe how the molecule relates to your career. Draw the molecule's structure and explain how it determines the properties that make this molecule important." Along with other early arrivals, Scott and Jennifer settle into their chairs and begin to discuss their career plans and the molecules important in their chosen fields.

Scott, an environmental science major, is interested in fuel-efficient engines, so he thinks octane should be his choice. He knows that octane is a standard for rating grades of gasoline, but it isn't until he checks his textbook that he learns that octane, C_8H_{18}, has several possible structures, including



As Scott studies the pictures, he realizes that the octane structures differ only in the way the carbon atoms attach to one another. Because of the large number of atoms, Scott learns that octane molecules have London dispersion forces strong enough to make them liquids at room temperature. (The respective boiling points of the molecules shown are 125.6°C, 117.6°C, and 99.3°C.) Scott also finds that only one of the structures—the last molecule shown—serves as the standard for the octane rating of gasoline. It burns more smoothly than the other molecules.

These octane compounds all have the same formula and the same mass. Why would they have different boiling points?

Jennifer hopes to have a career in dentistry, so she thinks about mouthwashes and a common ingredient in them: menthol, $C_{10}H_{20}O$. The menthol molecule, she finds, has the following structure:



What group is present in menthol that is not found in octane?

Menthol is also used in local anesthetics, such as cough drops, ointments for muscle aches, and anti-itch creams. It is used in aftershave lotions and skin fresheners as well. In all of these applications, menthol gives the sensation of cooling. Jennifer guesses that receptors on cell membranes in the mouth, throat, and skin might accept the menthol molecule and generate nerve impulses as a result. If this

were true, then the size, shape, and polarity of the menthol molecule would have to be just right for it to interact with a receptor, she reasons. Further reading proves she's right. Although the mechanisms aren't totally understood, research evidence suggests that menthol tricks the temperature receptors into signaling the brain that something cold has come into contact with them.

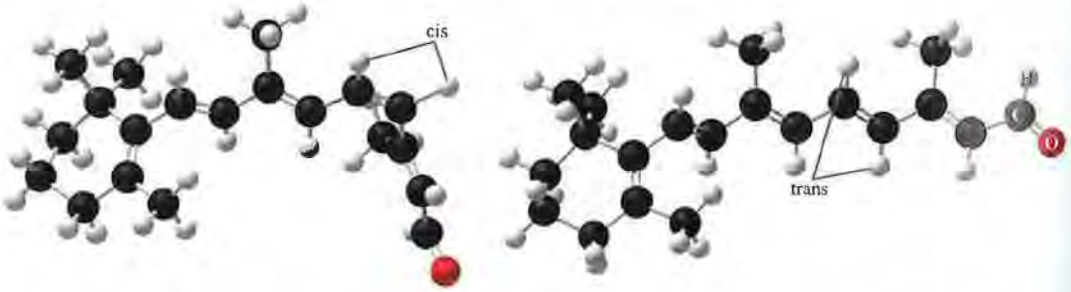
Omar plans to become a pharmacist. He picks aspirin, $C_9H_8O_4$, as his molecule. He finds the following structure for aspirin, also called acetylsalicylic acid:



Aspirin blocks the action of an enzyme called cyclooxygenase, Omar learns. The enzyme catalyzes the production and release of prostaglandins. Prostaglandins form whenever tissue is injured, and they send a nerve signal that the brain interprets as pain. Blocking the production of prostaglandins thus decreases the pain felt from an injury, although it does nothing to cure the cause of the pain. Prostaglandins also cause blood platelets to stick together and form clots. Blocking the production of prostaglandins helps to prevent blood clots, so aspirin is often taken to treat or prevent strokes and heart attacks. Omar suspects that aspirin inhibits the functioning of the enzyme by fitting into a cavity in the enzyme structure, blocking substrate molecules from attaching to it.

Would you expect the structure of other painkillers, such as acetaminophen (one brand name is Tylenol), to be similar to that of aspirin?

Maya hopes to become an ophthalmologist. She picks retinal, $C_{20}H_{28}O$, as her molecule since it is important to vision. She finds two different structures for this molecule, called *cis*-retinal (left) and *trans*-retinal (right):



The *cis*-retinal molecule has two hydrogen atoms on the same side of one of the carbon-carbon double bonds, while *trans*-retinal has the two hydrogen atoms on opposite sides of this double bond. This difference in structure causes a difference in shape, as seen in the models. The $-CHO$ group on the end of the molecule is quite reactive. It allows a *cis*-retinal molecule to bond to a molecule of opsin, a protein in the rods and cones of the eye. The result is a molecule called rhodopsin (Figure 16.1). The shape of *cis*-retinal allows it to fit into a cavity in the opsin molecule. Visible light converts *cis*-retinal to *trans*-retinal, accompanied by a change from a more bent to a more linear molecular structure. Because it no longer fits well into the opsin molecule cavity, *trans*-retinal detaches from the protein, sending an electrical signal along the optic nerve to the brain. The *trans*-retinal then interacts with an enzyme and changes back to *cis*-retinal, which can reattach to an opsin molecule.

The behavior of retinal is unusual. Hydrocarbons with double bonds do not normally convert from cis to trans forms upon exposure to light because the conversion requires one of the bonds to break.

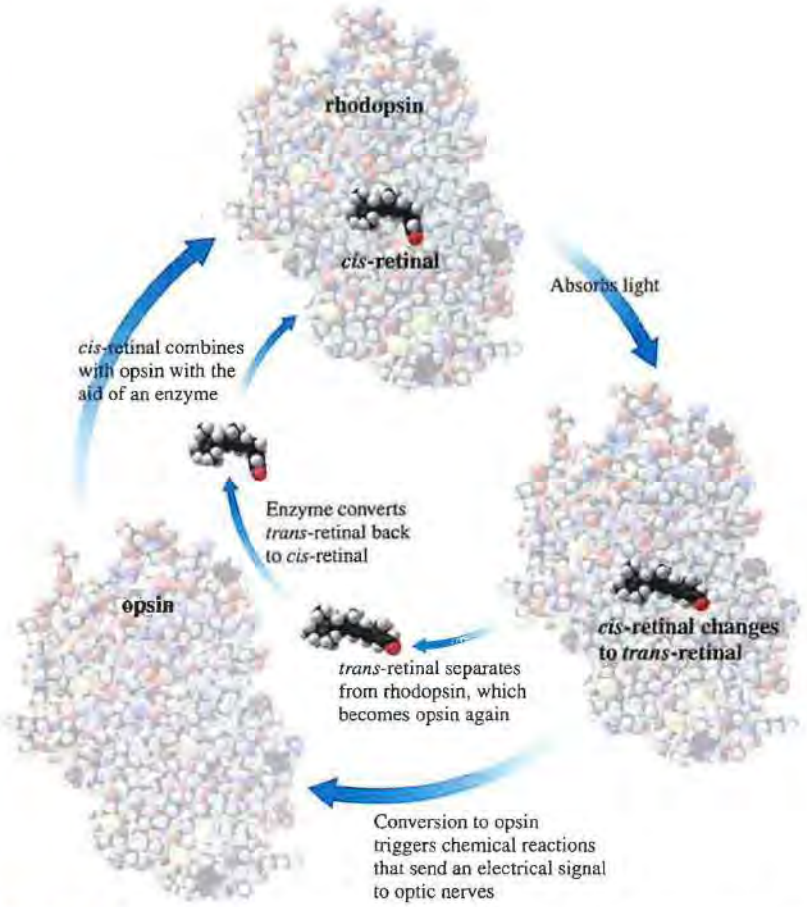


FIGURE 16.1 Vision depends on the interconversion of *cis*-retinal and *trans*-retinal. The absorption of light causes bond breakage, allowing the interconversion to occur.

Aisha, a criminal justice major, is concerned about law enforcement and drug abuse, so she decides to research the addictive central nervous system stimulant amphetamine, $C_9H_{13}N$. She finds this structure for the molecule:



How are the structures of aspirin and amphetamine alike and different?

The structure and shape of amphetamine are similar to that of norepinephrine, a natural neurotransmitter that carries nerve impulses across the gap between neurons in the brain. Norepinephrine is normally stored inside the ends of neurons before it is released in appropriate amounts. Because of its structure, amphetamine can displace norepinephrine from its storage vesicles. As a result, too many norepinephrine molecules are released into the gap, and emotional responses are exaggerated.

What do all these molecules have in common? Did you notice that they all contain carbon? Why is this important? If you had this assignment, what molecule would you select? Would you choose a carbon compound? Try to find a molecule that would meet the criteria outlined by the professor in the first paragraph of this

Some other carbon-containing molecules that are related to specific careers are

lactic acid (in weary muscles)	sports medicine and physical therapy
linoleic acid (in linseed oil)	painting
methyl methacrylate (Plexiglas)	sculpture
glycine (an essential amino acid)	nutrition
glucose	emergency medical technician
saturated fats	food science
diethyl ether	anesthesiology

section and carry out the assignment. In this chapter, we'll look at compounds of carbon and hydrogen, and we will explore the field of *organic chemistry.*

Questions for Consideration

16.1 How can we represent organic molecules to illustrate their similarities and differences?
16.2 What are hydrocarbons, and where do they come from?
16.3 What are acyclic hydrocarbons, and how do they react?
16.4 What are cyclic hydrocarbons, and how do they react?
16.5 How do alcohols and ethers differ from the alkanes?
16.6 How are aldehydes and ketones alike and different?
16.7 How do carboxylic acids and esters differ from other organic compounds?
16.8 What is unique about amines?
16.9 How do we use the IUPAC system to name organic compounds?

16.1 | REPRESENTATIONS OF ORGANIC MOLECULES

The compounds that Scott, Jennifer, and the other students chose for the assignment all contain carbon. In combination with a few other elements (primarily H, O, N, and S), carbon forms several million known compounds, far exceeding those of any other element except hydrogen. The most extensive class of carbon compounds consists of *organic compounds,* which contain carbon, hydrogen, and sometimes other elements. The study of organic compounds is called **organic chemistry**. The term *organic* relates to the importance of carbon compounds in living systems. Organic compounds need not come from living things, but they are vital to all life processes.

The common definition of an organic compound is any compound that contains carbon, but this includes carbonates (CO_3^{2-}), cyanides (CN^-), and other ionic compounds that are usually considered inorganic. Organic chemistry is better defined as the study of carbon compounds in which carbon atoms are bonded to one another and to hydrogen atoms as well, and compounds that are derivatives of these hydrocarbons.

Why is carbon found in so many molecules and in such a variety of structural forms? The answer is found, in part, in the way that carbon bonds to itself and to other elements. A carbon atom can form single bonds, double bonds, or triple bonds in various combinations—for example,

- four single bonds, as in C_2H_6 
- two single bonds and one double bond, as in C_2H_4 
- two double bonds, as for the central atom in C_3H_4 
- one single bond and one triple bond, as in C_2H_2 

In each case carbon has a total of four bonds. In contrast, hydrogen and fluorine typically form one bond, oxygen forms two bonds, and nitrogen forms three. The greater number of bonds found with carbon gives more possibilities for forming a diverse set of compounds.

A carbon compound that contains only carbon and hydrogen is called a **hydrocarbon**. In hydrocarbons, carbon has the ability to bond to other carbon atoms to form long chains or rings of carbon atoms with hydrogen atoms attached. Different branching patterns are possible in many hydrocarbon molecules. Two compounds having the same composition and formula, but with a different arrangement of atoms in their molecular structures, are **isomers** of one another.

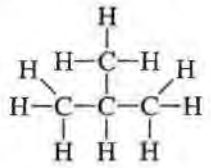
Carbon can bond to another nonmetal, such as nitrogen and oxygen, in a few, well-defined ways, giving a characteristic grouping of atoms called a functional group. A **functional group** is a reactive part of a molecule that undergoes charac-

teristic reactions. For example, when a hydroxyl group (−OH) is present instead of a hydrogen atom in propane, C_3H_8, we have propyl alcohol or propanol, C_3H_7OH. The functional group in this case is −OH, which is characteristic of alcohols.

When the functional group undergoes a reaction, the remainder of the molecule usually remains intact. In reactions, then, the hydrocarbon chain to which the functional group is attached is often unchanged, so for simplicity we often abbreviate the hydrocarbon part of the molecule as R−, and show it attached to the functional group. If two hydrocarbon groups are present, the second can be designated R′−. The representation of any alcohol, for example, is R−OH. The shorthand R allows us to ignore the hydrocarbon chain and concentrate on the part of the molecule that gives it many of its characteristic properties.

A variety of related molecules with different functional groups is shown in Table 16.1. How are these structures related to propane? How can substituting, adding, or removing atoms or groups of atoms create these structures?

Ball-and-stick models are a good way to represent the structure of organic molecules, but other symbols and models have their advantages. The hydrocarbon fuel propane, for example, has the simple *molecular formula* C_3H_8. It has three carbon atoms linked in a chain, with hydrogen atoms bonded to each carbon atom in sufficient numbers to give each carbon atom four bonds: $CH_3CH_2CH_3$. This *structural formula* gives more information about propane than does the molecular formula, because it shows the order in which the atoms are bonded. Butane, another fuel, has the molecular formula C_4H_{10}. It can be represented by a structural formula $CH_3CH_2CH_2CH_3$, or a *condensed structural formula*, $CH_3(CH_2)_2CH_3$, in which identical connected units are grouped. Butane can also exist as an isomer in which one carbon atom is bonded to three $-CH_3$ groups and one hydrogen atom. The condensed structural formula for the isomer is $CH_3CH(CH_3)CH_3$, where the part of the formula in parentheses is understood to be attached to the previous carbon atom:



Another way to represent organic molecules is a *line structure*, which represents connected carbon atoms with a line. The hydrogen atoms bonded to carbon atoms are not shown, because their presence can be deduced from the fact that carbon must form four bonds. Any other atom is specified by its element symbol. In a line structure, the intersection of two lines or any unconnected end of a line represents a carbon atom. Thus, the octane molecules that Scott is interested in can be represented in the following ways:

- Ball-and-stick model: 
- Molecular formula: C_8H_{18}
- Structural formula: $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
- Condensed structural formula: $CH_3(CH_2)_6CH_3$
- Line structure: 

The ball-and-stick model, condensed structural formula, and line structure for the second isomer of octane that Scott found are



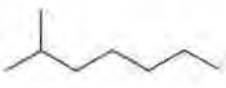
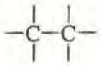
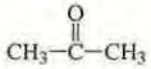
$(CH_3)_2CH(CH_2)_4CH_3$ 

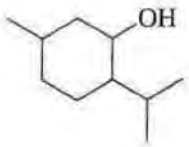
TABLE 16.1 | Functional Groups in Organic Molecules

Type	Functional Group	Example	Example Structure	Example Name	Example Property
alkane		$CH_3CH_2CH_3$		propane	heating fuel for camping stoves
alkene		$CH_3CH{=}CH_2$		propene	used to form polypropylene, a stiff plastic
alkyne	$-C{\equiv}C-$	$CH_3C{\equiv}CH$		propyne	explosive in air
alcohol	$-OH$	$CH_3CH_2CH_2OH$		propanol	alcoholic and slightly stupefying odor
ether	$-O-$	$CH_3CH_2CH_2OCH_3$		methyl propyl ether	anesthetic
ketone				acetone	good solvent; fingernail polish remover
aldehyde		CH_3CH_2CHO		propanal	suffocating odor
carboxylic acid		$CH_3CH_2CO_2H$		propanoic acid	odor and flavor of Swiss cheese
ester		$CH_3CH_2CO_2CH_3$		methyl propanoate	flavoring agent; fruity rum odor
amine		$CH_3CH_2CH_2NH_2$		propylamine	strong ammonia odor

The line structure is a convenient way to show the difference in isomers. In these examples, the first isomer is a straight chain of eight carbon atoms, while in the second, one of the carbons is bonded to the second carbon atom in a seven-carbon chain.

Line structures also help us focus on the functional group where any chemical reaction is likely to occur. The line structure for Jennifer's menthol molecule is





The molecule is a ring, or cycle, of six connected carbon atoms. Attached to this ring are a carbon atom, an −OH group (after skipping one carbon atom in the ring), and then a three-carbon-atom chain connected by its center carbon atom. The menthol molecule has the −OH functional group that is characteristic of alcohols, so it has physical and chemical properties similar to that of other alcohols.

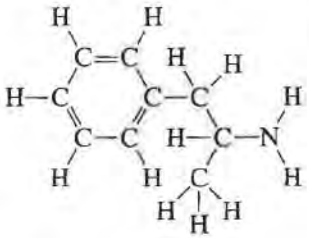
EXAMPLE 16.1 | Drawing Structural Formulas and Line Structures

Aisha studied the amphetamine molecule, using the following ball-and-stick model. Draw both a structural formula and a line structure for this molecule:

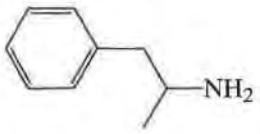


Solution:

Examine the ball-and-stick model of this molecule. The structural formula must represent each carbon-carbon bond and include the hydrogen or other atoms attached to each carbon. Since this is not a simple chain structure, we use lines to represent bonds and show how the atoms are connected:



The line structure can then be based on this structural formula, but with the carbon-hydrogen bonds and the symbols for carbon removed:



For clarity we leave in the element symbols for the NH_2 group that is characteristic of amines.

Practice Problem 16.1
Omar studied the aspirin molecule with the following ball-and-stick model:

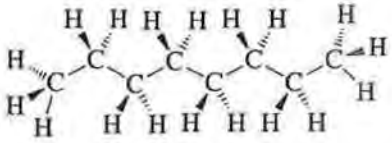


Draw a structural formula and a line structure for this molecule.

Further Practice: Questions 16.7 and 16.8 at the end of the chapter

One additional representation of organic molecules is a *perspective drawing*, which can be used instead of a ball-and-stick model when visualizing the three-dimensional structure of a molecule. In a perspective drawing, solid lines represent bonds that are in the plane of the paper. Dashed wedges represent bonds that extend back behind the plane of the paper. Solid wedges represent bonds that project in front of the plane of the paper. To see how this works, compare the perspective drawing of octane to the ball-and-stick model:





16.2 | HYDROCARBONS

Classes of Hydrocarbons

Two major kinds of hydrocarbons, the *aliphatics* and the *aromatics*, were described in Section 8.3 and outlined in Figure 8.20.

Hydrocarbons are a major class of organic compounds, and they form the basic structural components of most other classes. Hydrocarbons differ in the type of carbon-carbon bonds they contain and in whether the carbon atoms form open chains or closed rings. The hydrocarbons can be classified as shown in Figure 16.2.

In organic chemistry *saturated* and *unsaturated* have different meanings than when describing the concentration of solutions, but the two uses are related. In each case *unsaturated* means some "holding capacity" remains—either to accept more hydrogen atoms (organic molecules) or to dissolve more solute (in solutions). *Saturated* means that all capacity to hold more has been used.

Acyclic (Open Chain) Hydrocarbons The first main categories shown in Figure 16.2 are derived from the shape of the molecule. A hydrocarbon molecule that forms an open chain is an **acyclic hydrocarbon**. In such molecules the bonding electrons are *localized;* the electrons are confined to two adjacent atoms. A hydrocarbon with localized bonds is an **aliphatic hydrocarbon**.

A **saturated hydrocarbon** contains only single bonds between the carbon atoms. All of the carbon atoms are said to be saturated because they are attached to four other atoms—the maximum number possible for carbon. An example is ethane, H_3C-CH_3, which, like all the other compounds in this category, is an alkane. An *alkane* is any straight-chain hydrocarbon containing only carbon-carbon single bonds.

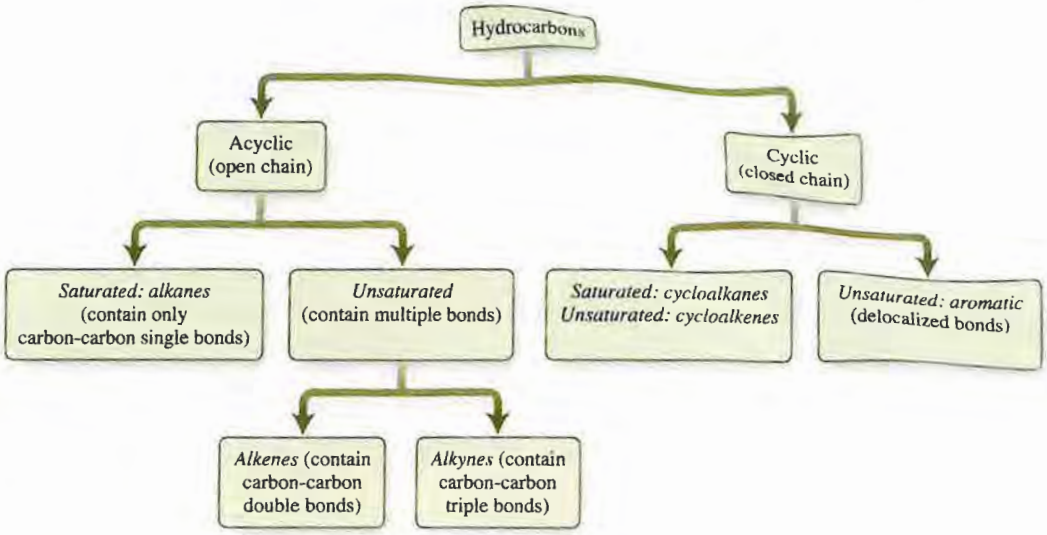


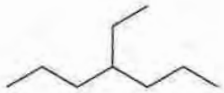
FIGURE 16.2 Classification of the hydrocarbons.

An **unsaturated hydrocarbon** contains one or more double or triple bonds between carbon atoms. Such molecules are considered unsaturated because two or more of the carbon atoms do not have four other atoms bonded to them. The double-bonded, unsaturated hydrocarbons are the *alkenes,* such as $H_2C{=}CH_2$ (ethene, also called ethylene). The triple-bonded, unsaturated hydrocarbons are the *alkynes,* such as $HC{\equiv}CH$ (ethyne, also called acetylene).

All of the acyclic, aliphatic hydrocarbons form open chains of carbon atoms, such as propane, $H_3C{-}CH_2{-}CH_3$. The chains may bond to side chains, or branches, such as in 4-ethylheptane:



$$\begin{array}{l} \quad\quad\quad H_2C{-}CH_3 \\ \quad\quad\quad\quad | \\ H_3C{-}CH_2{-}CH_2{-}CH{-}CH_2{-}CH_2{-}CH_3 \end{array}$$



We'll examine the three categories of acyclic hydrocarbons in more detail in Section 16.3.

Cyclic (Closed Chain) Hydrocarbons In contrast to the open chain of an aliphatic compound, a **cyclic hydrocarbon** forms a closed chain or ring of carbon atoms, such as in cyclobutane:



$$\begin{array}{l} H_2C{-}CH_2 \\ \quad | \quad\quad | \\ H_2C{-}CH_2 \end{array}$$

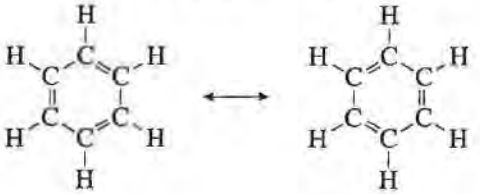


Cyclobutane, C_4H_8, is a colorless gas that burns with a luminous flame.

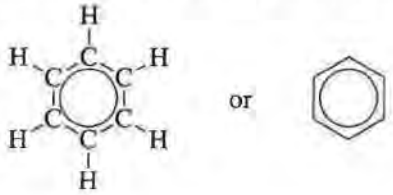
This is an example of a *cycloalkane:* a single-bonded chain that appears to have bent around on itself and attached at its ends. Cycloalkanes, like alkanes, are single-bonded and saturated.

In contrast, an **aromatic hydrocarbon** is a cyclic molecule containing a ring of carbon atoms attached to each other and is represented by resonance structures containing alternating single and double bonds. The double-bond electrons in these

compounds are *delocalized*, as explained by resonance (review Section 8.3). For example, a six-carbon ring such as in benzene, C_6H_6, has two possible resonance structures:



However, neither of these drawings represents the actual structure. Instead, the double-bond electrons in benzene are delocalized or spread out around the entire ring structure, giving the molecule bonds of equal length and—unlike ordinary double bonds—a tendency toward low reactivity. For these reasons the structure is more accurately represented with a circle inside the ring to indicate the delocalized double bonds:



We'll consider aromatic hydrocarbons in more detail in Section 16.4.

ANIMATION: Petroleum Refinery

Petroleum is a primary source of carbon-containing compounds in plastics and other products. Should we continue to use most of our petroleum as fuel?

Petroleum

Along with coal, petroleum is the primary source of hydrocarbons. Their major use is energy production, but around 10% of the coal and petroleum recovered from Earth is used to make such useful materials and products as plastics, drugs, foods, and synthetic fibers.

Petroleum is usually found naturally in one of two forms, either natural gas or crude oil. Natural gas consists mainly of methane (CH_4), with smaller amounts of ethane (C_2H_6), propane (C_3H_8), and butane (C_4H_{10}). All are colorless, odorless, flammable gases. Natural gas is used primarily for heating and cooking. Crude oil is a dark, thick, odoriferous liquid composed of about 95% hydrocarbons. It must be refined before it can be used. Refining involves a few basic steps. The first is *fractional distillation*, or the separation of the petroleum into fractions having different ranges of boiling points. These fractions are either used directly, purified further, or converted into products that are more desirable. The fractions typically obtained from petroleum are summarized in Table 16.2.

Because natural gas has no odor, traces of methyl mercaptan, CH_3SH, are added to commercial supplies so gas leaks in homes and businesses can be detected. It has the odor of rotten eggs.

Most petroleum is used to make gasoline, a mixture of hydrocarbons of which only one is octane, the molecule Scott chose for study. Given our reliance on gasoline as a fuel for transportation, the fraction of petroleum that can be directly separated as gasoline is insufficient to meet demand, because more than half of the hydrocarbons in petroleum have more than 13 carbon atoms, the longest chain found in the gasoline fraction. To obtain larger amounts of gasoline, the process of *catalytic cracking* uses heat and a catalyst to break down the higher-boiling "gas oil" fraction. For example,

Recall from Chapter 12 that a catalyst acts to increase the rate of a reaction.

$$C_{14}H_{30} \xrightarrow{\text{heat, catalyst}} C_6H_{12} + C_8H_{18}$$

Note the wide range of boiling points in Table 16.2. Why do hydrocarbons boil at such different temperatures? Hydrocarbons differ in their physical characteristics as the number and arrangement of carbon atoms change. The strengths of the intermolecular forces that attract the molecules to one another account for observed differences in properties. For example, as the number of carbon atoms increases, the physical state of the hydrocarbon changes from gas to liquid to solid. What happens

Hydrocarbons containing from 5 to 18 carbon atoms are generally liquid at ordinary environmental temperatures. A number of everyday hydrocarbon products are liquids. Gasoline is made up of the lighter hydrocarbons. Kerosene contains slightly heavier hydrocarbons, and lubricating oil contains even heavier hydrocarbons.

TABLE 16.2 | Petroleum Fractions

Fraction	Boiling Range (°C)	Composition Range	Uses
natural gas	< 20	CH_4 to C_4H_{10}	fuel, petrochemicals
petroleum ether	20–60	C_5H_{12}, C_6H_{14}	solvent
ligroin, or naphtha	60–100	C_6H_{14}, C_7H_{16}	solvent, raw material
gasoline	40–220	C_4H_{10} to $C_{13}H_{28}$, mostly C_6H_{14} to C_8H_{18}	motor fuel
kerosene	175–325	C_8H_{16} to $C_{14}H_{30}$	heating fuel and jet fuel
gas oil	> 275	$C_{12}H_{26}$ to $C_{18}H_{38}$	diesel fuel and heating fuel
lubricating oils and greases	high: viscous liquids	> $C_{18}H_{38}$	lubrication
paraffin	high: melting point 50–60	$C_{23}H_{48}$ to $C_{29}H_{60}$	wax products
asphalt or petroleum coke	high: solid residue		roofing, paving, fuel, reducing agent

to the intermolecular forces? As the molecules get larger, they have greater surface area. Their London dispersion forces increase, so they interact more strongly and their boiling points increase.

16.3 | ACYCLIC HYDROCARBONS

The simplest organic compounds are the open-chain, acyclic hydrocarbons: the alkanes, alkenes, and alkynes. The chains may be straight or branched. The alkanes are single-bonded and saturated. The alkenes and alkynes are unsaturated.

Alkanes

An **alkane** is a saturated hydrocarbon with the general formula C_nH_{2n+2}. The alkanes are obtained either directly from petroleum or indirectly by *cracking,* which breaks larger hydrocarbons into smaller ones. When the value of n is 1 through 4, the alkanes are gases under normal conditions. When n equals 5 through 18, they are liquids; and when n is greater than 18, they are solids. The melting points and boiling points increase as the number of carbon atoms increases. The increase results from the greater surface area and stronger intermolecular forces that accompany the larger molecular size. A graph of the change in boiling point as the length of the carbon chain increases is shown in Figure 16.3.

Consider some examples of alkanes that are consistent with the general formula C_nH_{2n+2}. If $n = 1$, $2n + 2 = 4$, and we have CH_4. If $n = 3$, $2n + 2 = 8$, and we have C_3H_8 (propane).

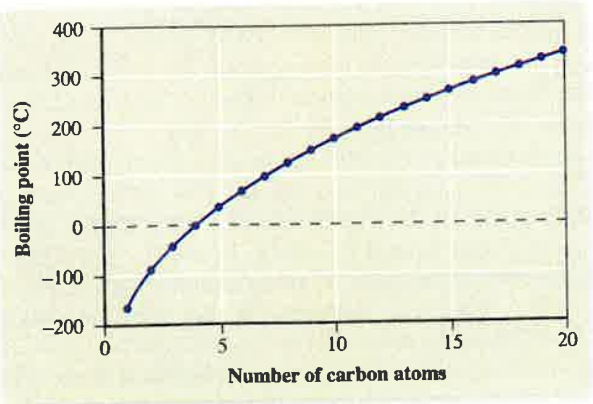


FIGURE 16.3 The boiling points of straight-chain alkanes increase as the number of carbon atoms in the chain increases. This is due to the increase in intermolecular forces.

Straight-chain alkanes are not really straight, since each carbon is at the center of a tetrahedron. A straight-chain alkane thus has a zigzag structure. For this reason, some chemists call these alkanes continuous-chain alkanes or unbranched-chain alkanes.

As the size of the molecule affects its intermolecular forces, so too does the shape. Straight-chain molecules can have more surface contact with one another than branched-chain molecules (Figure 16.4). The increased surface contact leads to increased London dispersion forces, with temporary dipoles attracting one another. When the molecules are more highly attracted, more energy has to be added to separate the molecules as they escape into the gas phase during boiling.

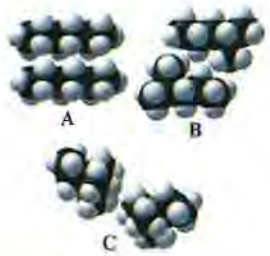


FIGURE 16.4 Straight-chain alkanes (A) have more surface contact between molecules than branched-chain alkanes (B and C) containing the same number of atoms. Thus, straight-chain alkanes experience greater London dispersion forces than do branched-chain alkanes, resulting in higher boiling points. The boiling points of these alkanes, all with the formula C_6H_{14}, are (A) 69.0°C, (B) 60.3°C, and (C) 49.7°C.

EXAMPLE 16.2 | Intermolecular Forces in Alkanes

Scott found boiling points for two hydrocarbons having eight carbon atoms, represented here by their line structures. Explain the difference in boiling points of these two molecules:



bp = 125.6°C bp = 117.6°C

Solution:

Because these molecules contain only carbon and hydrogen, they are nonpolar and they experience only London dispersion forces. They have the same molecular formula, so they have the same number of electrons. The greater surface contact between the linear molecules (left) provides more chances for temporary dipoles to form and interact. Branched-chain molecules (right) have less surface contact than straight-chain molecules, and they don't fit next to one another as well. Thus, the branched molecule has the smaller intermolecular forces. It boils at a lower temperature because overcoming its intermolecular forces requires a lesser input of energy.

Practice Problem 16.2

Scott also found a boiling point for another hydrocarbon having eight carbon atoms, represented by this line structure:



bp = 99.3°C

Explain why this boiling point is lower than the other two given in the example.

Further Practice: Questions 16.23 and 16.24 at the end of the chapter

Alkane Structure and Nomenclature The alkanes are named with a prefix denoting the number of carbon atoms in the longest chain followed by the ending *-ane*. The four shorter alkanes—methane (CH_4), ethane (C_2H_6), propane (C_3H_8), and butane (C_4H_{10})—have special names, while the longer alkanes use prefixes derived from the Greek or Latin numbers, such as octane for C_8H_{18}. The names of the first 10 alkanes are listed in Table 16.3.

Each successive alkane in the table can be considered to be derived by insertion of one additional $-CH_2-$ group between the first carbon atom and one of its attached hydrogen atoms in the smaller alkane. For example, insertion of a CH_2 group into methane gives ethane, CH_3-CH_2-H or CH_3-CH_3. The third alkane is propane (C_3H_8) which has a second CH_2 group inserted to give $CH_3-CH_2-CH_2-H$ or $CH_3-CH_2-CH_3$. The next molecule in the series is butane (C_4H_{10} or $CH_3-CH_2-CH_2-CH_2-H$ or $CH_3-CH_2-CH_2-CH_3$).

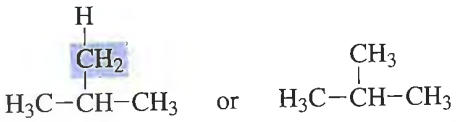
ANIMATION: Isomers

Alkanes in which the carbon atoms form a continuous chain, with no branches, are called *normal alkanes* or *straight-chain alkanes*. The alternate arrangements are

TABLE 16.3 | Names of Simple Alkanes

Alkane Name	Formula (C_nH_{2n+2})
methane	CH_4
ethane	C_2H_6
propane	C_3H_8
butane	C_4H_{10}
pentane	C_5H_{12}
hexane	C_6H_{14}
heptane	C_7H_{16}
octane	C_8H_{18}
nonane	C_9H_{20}
decane	$C_{10}H_{22}$

called *branched-chain alkanes* and are isomers of the straight-chain compounds. Thus—unlike the first three alkanes—butane has two different structures, or isomers. The first is normal butane (*n*-butane or just butane); it is an unbranched chain of four carbon atoms. The second is isobutane, a chain of three carbon atoms with the added CH_2 group inserted at a center carbon:



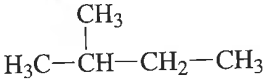
The number of isomers increases rapidly as the number of carbon atoms increases. Pentane, C_5H_{12}, has three isomers, $C_{10}H_{22}$ has 75 isomers, and $C_{20}H_{42}$ has 366,319 isomers. Because of the large number of possible isomers, alkanes are named in accordance with a systematic nomenclature established by the International Union of Pure and Applied Chemistry (IUPAC). If common names were applied to all of these isomers, chemists would never be able to communicate efficiently. Consider the molecules and their IUPAC names given in Table 16.4. Can you determine the rules used for naming these compounds?

Did you recognize how to name side chains? Did you figure out how to use a numbering system to indicate position? If you need help with the rules, consult Section 16.9. Try to apply your rules to name an alkane with one branch on the chain.

In a molecule containing more than one of a group, prefixes indicate multiples. For example, *di-* means two and *tri-* means three.

EXAMPLE 16.3 | Names of Alkanes with a Single Branch

Name the following compound—known commonly as isopentane—according to the IUPAC nomenclature rules:



Solution:

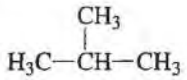
Step 1: The longest continuous carbon chain has four carbons, so the name is based on butane.

Step 2: The carbons in this chain are numbered from the end closer to a branch. The only branch occurs at position 2.

Step 3: The branch occurring at position 2 is CH_3, or methyl, named for the alkane it is derived from, methane.

Step 4: The name is 2-methylbutane.

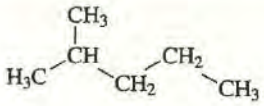
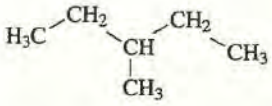
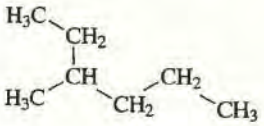
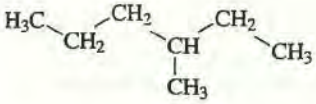
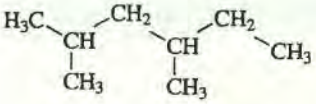
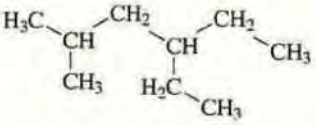
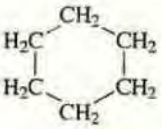
Practice Problem 16.3
Name the following compound—known commonly as isobutane—according to the IUPAC nomenclature rules:



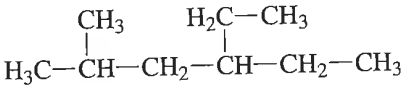
Further Practice: Questions 16.31 and 16.32 at the end of the chapter

When naming the branches on alkanes, look for a group derived from an alkane by removal of one hydrogen atom, with a general formula C_nH_{2n+1}. Such a group is an **alkyl group.** The names of the alkyl groups are similar to those of the alkanes (given in Table 16.3), except that the *–ane* ending is replaced by *–yl*. Thus, $-CH_3$ is methyl and $-C_2H_5$ is ethyl.

TABLE 16.4 | Formulas and Names of Some Alkanes

Structural Formula	Name
	2-methylpentane
	3-methylpentane
	3-methylhexane
	3-methylhexane
	2,4-dimethylhexane
	4-ethyl-2-methylhexane
	cyclohexane

If an alkane contains more than one branch, each is named separately and in alphabetical order. Consider the following molecule:



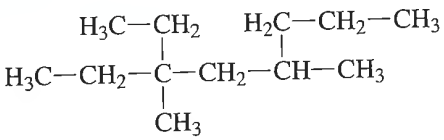
The longest chain is hexane. A methyl ($-CH_3$) branch occurs at position 2 and an ethyl ($-C_2H_5$) branch occurs at position 4. The numbering starts from the left end of the molecule to give lower numbers to the branch positions. The ethyl group is named first since it occurs first alphabetically. The molecule is named 4-ethyl-2-methylhexane.

Notice that the third and fourth examples listed in Table 16.4 have the same name. Do these structures represent the same compound? Which end of each structure is used to start numbering the chain? Convince yourself that these two structures represent the same compound. The IUPAC system ensures that each compound has a unique name. The following examples will help you learn the basic rules for the IUPAC system.

Now try to name another alkane with more than one branch. If you need help with the rules, consult Section 16.9.

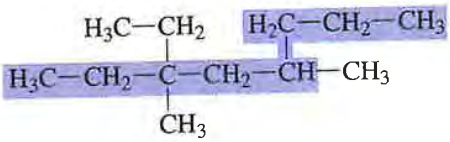
EXAMPLE 16.4 | Names of Alkanes with Multiple Branches

Name the isomer of $C_{12}H_{26}$ that has the following structure:



Solution:

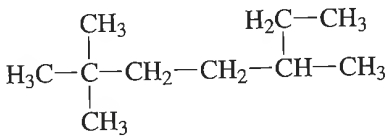
The longest continuous chain has eight carbons, as shown by the highlighted part of the structure, so the compound is named as a derivative of octane.



There are two methyl groups, one at carbon 3 and one at carbon 5, and one ethyl group at carbon 3. The ethyl group should be named first, because it occurs first alphabetically. The methyl groups can be named together, using the carbon numbers to indicate their positions. The full name is thus 3-ethyl-3, 5-dimethyloctane.

Practice Problem 16.4

Name the isomer of $C_{10}H_{22}$ that has this structure:



Further Practice: Questions 16.33 and 16.34 at the end of the chapter

If you are given a name, can you figure out the structure of the compound? You can use the rules in reverse to yield the structural formula from the IUPAC name.

The octane rating is based on the relative difference in the ability of a compound in gasoline to cause knocking in an engine. The highly branched structure of 2,2,4-trimethylpentane gives it the characteristic of burning smoothly with little knocking in an engine. This compound has an octane rating of 100. In contrast, the unbranched structure of *n*-heptane causes an engine to knock badly, and *n*-heptane has been assigned an octane rating of 0. Gasoline is a mixture of compounds of differing degrees of branching, so the octane rating of most gasoline falls somewhere in between this range and is usually around 90.

EXAMPLE 16.5 | Structural Formulas from Names of Alkanes

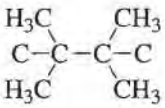
Although Scott found three isomers having the chemical formula C_8H_{18}, there are actually more. One is 2,2,3,3-tetramethylbutane. Draw a structural formula for this molecule.

Solution:

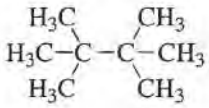
The molecule name contains butane, so the longest chain must contain four carbon atoms:



Of the four methyl groups ($-CH_3$), two are attached to carbon 2. The other two are attached to carbon 3.



Finally we add enough hydrogen atoms to give four bonds to each carbon atom:



Practice Problem 16.5

Another isomer of C_8H_{18} is 2,2,4-trimethylpentane. This compound in gasoline causes little engine knocking and has the hightest possible octane rating. Draw its structural formula.

Further Practice: Questions 16.35 and 16.36 at the end of the chapter



A



B

FIGURE 16.5 Hexane does not react with brown bromine unless exposed to bright light. (A) When hexane and bromine are mixed in the absence of light, the mixture remains brown due to the presence of unreacted bromine. (B) In the presence of light, a reaction occurs in which the product is colorless.

Reactions of Alkanes Alkanes don't readily undergo chemical reactions, because both carbon-hydrogen and carbon-carbon single bonds are strong. Supplying heat or light, however, provides the energy needed to get alkanes to react. For example, in petroleum refining, heating in the presence of a catalyst cracks or breaks large alkanes into smaller molecules. Such reactions produce shorter-chain alkanes, cycloalkanes, alkenes, and molecular hydrogen.

Alkanes undergo oxidation in exothermic *combustion* reactions, in which they act as fuels and burn when ignited in the presence of oxygen gas to form carbon monoxide or carbon dioxide and water:

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

In the presence of sunlight or ultraviolet radiation, alkanes will react with halogens. As shown in Figure 16.5, light provides the energy to break the halogen-halogen bonds and initiate the reaction:

$$C_6H_{14} + Br_2 \xrightarrow{\text{light}} C_6H_{13}Br + HBr$$

EXAMPLE 16.6 | Reactions of Alkanes

Predict the products of the reaction of propane with bromine in the dark at room temperature.

Solution:

Alkanes do not react with bromine unless subjected to heat, sunlight, or ultraviolet radiation. Thus, there would be no reaction under the conditions given.

Practice Problem 16.6

Predict the products of the reaction of butane with bromine in sunlight.

Further Practice: Questions 16.37 and 16.38 at the end of the chapter

Alkenes and Alkynes

Maya's molecule, retinal, contains two types of functional groups: a carbon-carbon double bond and a −CHO group. Later we'll look at the properties of the −CHO group. Here we will be concerned about the effect that double bonds have on the properties of hydrocarbons. In this section, we consider two such classes of hydrocarbons: alkenes, containing double bonds, and alkynes, containing triple bonds.

Alkene Structure and Nomenclature **Alkenes** contain at least one C=C double bond, and with just one double bond have the general formula C_nH_{2n}. They are generally obtained from petroleum by the cracking of alkanes, using heat and a catalyst. The simplest such reaction is

$$CH_3{-}CH_3 \xrightarrow{\text{heat, catalyst}} H_2C{=}CH_2 + H_2$$

The product is named ethene (IUPAC name), or ethylene (common name).

An alkene molecule may also have more than one double bond. For example, $H_2C{=}CH{-}CH{=}CH_2$, with double bonds at carbons 1 and 3, is 1,3-butadiene. Alkenes can also form *polymers*, molecules made of many repeating units. Polymers are used to make a variety of useful items including rubber tires.

Examine the structural formulas and names of some alkenes given in Table 16.5. Try to determine the rules for naming them.

Ethene seems to play a part in the ripening of fruit, so green fruit is often shipped under an atmosphere of ethene, so that ripening occurs in transit.

TABLE 16.5 | Formulas and Names of Some Alkenes

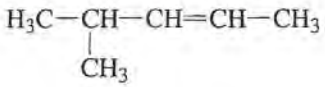
Structural Formula	Name
$H_2C{=}CH_2$	ethene
$H_2C{=}CH{-}CH_3$	propene
$H_2C{=}CH{-}CH_2{-}CH_3$	1-butene
$H_3C{-}CH_2{-}CH{=}CH_2$	1-butene
$H_3C{-}CH{=}CH{-}CH_3$	2-butene
$H_3C{-}\overset{\displaystyle CH_3}{\overset{\vert}{C}}{=}CH{-}CH_3$	2-methyl-2-butene
$H_2C{=}CH{-}\overset{\displaystyle CH_3}{\overset{\vert}{C}}H{-}CH_3$	3-methyl-1-butene
$H_2C{=}\overset{\displaystyle H_3C}{\overset{\vert}{C}}{-}\overset{\displaystyle CH_3}{\overset{\vert}{C}}H{-}CH_3$	2,3-dimethyl-1-butene

ANIMATION: Polymers

Did you come up with some rules? How do you name the main chain? How do you designate the positions of double bonds? How do you designate the positions and names of branches? If you need help, consult Section 16.9. Use your rules to name some alkenes.

EXAMPLE 16.7 | Names of Alkenes

Name this compound:

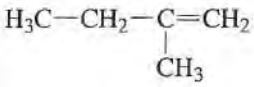


Solution:

The longest chain has five carbons, so the compound is a derivative of pentane, making the name pentene (because of the double bond). The carbon atoms are numbered from the end nearest the double bond. The position of the double bond is after the second carbon from the right. Counting from that end, the methyl side chain is attached to the fourth carbon atom. The name is 4-methyl-2-pentene.

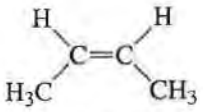
Practice Problem 16.7

Name this compound:

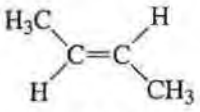


Further Practice: Questions 16.41 and 16.42 at the end of the chapter

In isomers of alkenes, rotation is restricted around the C=C double bond. Because of the restricted rotation, the relative positions of side chains attached to the carbons in the double bond are fixed. To have such *geometric isomers*, two different groups must attach to each of the carbon atoms in the double bond. The attached groups may be on the same side of the molecule (a *cis* isomer) or on opposite sides (a *trans* isomer). Thus, Maya found two different isomers of retinal—a cis form and a trans form—that differed in the positions of atoms attached to one of the C=C units in the molecule. A simpler example is 2-butene, with the two hydrogen atoms on the same or opposite sides of the double bond:



cis-2-butene



trans-2-butene

These two molecules have different properties because their structures are different. *Cis*-2-butene has a boiling point of 4°C, but that of *trans*-2-butene is 1°C. The cis isomer has stronger intermolecular forces because it is slightly polar, resulting in a higher boiling point.

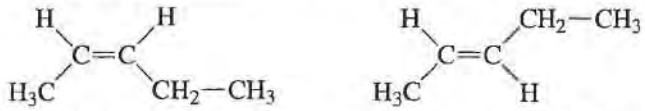
EXAMPLE 16.8 | Cis and Trans Isomers

Which of these molecules can have cis and trans isomers? Draw structures for any possible isomers.

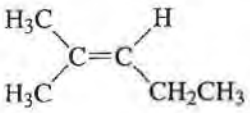
(a) $CH_3CH{=}CHCH_2CH_3$
(b) $(CH_3)_2C{=}CHCH_2CH_3$

Solution:

(a) Each double-bonded carbon has two groups attached to it, and the two groups in each case are different from each other. Thus, it is possible to have both cis and trans isomers:



(b) Each double-bonded carbon has two groups attached to it, but the first (left) carbon has identical groups. There is only one possible structure, so there are no cis and trans isomers.



Practice Problem 16.8

Which of these molecules can have cis and trans isomers? Draw structures for any possible isomers.

(a) $CH_3CH_2CH{=}CH_2$
(b) $CH_3CH_2CH{=}CHCH_2CH_3$

Further Practice: Questions 16.43 and 16.44 at the end of the chapter

Alkyne Structure and Nomenclature **Alkynes** contain a C≡C triple bond and have the general formula C_nH_{2n-2}. Alkynes are also called acetylenes, from the common name of the simplest alkyne, HC≡CH. Ethyne (IUPAC name) or acetylene (common name) is generally prepared from calcium carbide (Figure 16.6).

$$CaC_2(s) + 2H_2O(l) \longrightarrow HC{\equiv}CH(g) + Ca(OH)_2(s)$$

Alkynes are named with the same rules as alkenes, but using the ending *–yne* instead of *–ene*. Thus, $CH_3C{\equiv}CCH_3$ is 2-butyne and $HC{\equiv}CCH_2CH_3$ is 1-butyne. There are no cis and trans isomers for alkynes.



FIGURE 16.6 Before portable electric lamps became available, miners and cave explorers used lamps in which water dripped onto calcium carbide. The acetylene that was produced was burned to give light. Changing the rate of addition of water adjusted the size of the flame.

Reactions of Alkenes and Alkynes An *addition reaction* is one in which extra atoms or groups are added to a simpler hydrocarbon. Alkenes and alkynes typically undergo addition reactions across the double or triple bond:

$$CH_2{=}CH_2 + HBr \longrightarrow CH_3CH_2Br$$

Many of these reactions require a catalyst in order to occur at an appreciable rate. For example, if ethene is placed in an environment of H_2 gas, no reaction occurs. In the presence of a metal catalyst such as platinum, however, ethene converts to ethane:

$$CH_2{=}CH_2 + H_2 \xrightarrow{Pt} CH_3CH_3$$

Another important reaction is *polymerization*, in which small molecules containing double bonds combine together to form large molecules like plastics. The polymer is made of smaller, repeating units called *monomers*. The formation of polyethylene is an example of an *addition polymerization* reaction, in which monomers add to one another without losing any atoms. The joining of monomer units at the position of the monomer double bond forms the polymer, as shown here for the polymerization of ethylene to form polyethylene:

$$H_2C{=}CH_2 + H_2C{=}CH_2 + H_2C{=}CH_2 + \cdots \longrightarrow$$
$$-H_2C{-}CH_2{-}H_2C{-}CH_2{-}H_2C{-}CH_2{-}\cdots$$

This equation can be generalized for any number (n) of monomers as follows:

$$n H_2C{=}CH_2 \longrightarrow \text{-}(CH_2{-}CH_2)\text{-}_n$$



FIGURE 16.7 Many consumer products are made of polyethylene.

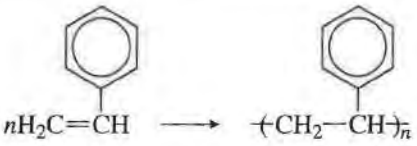
The groups shown in the parentheses are repeated in a chain n times, where n is a large number.

Polyethylene is a useful polymer. It can be heated to soften and shape it, and then cooled to a plastic substance with its original properties. That's why polyethylene bags can be heat-sealed. Polyethylene films hold in or keep out water vapor but allow the passage of oxygen. Polyethylene is a good electrical insulator. It is commonly used in soda bottles, milk cartons, produce bags, food wrap, wire coatings, and more (Figure 16.7).

Propylene is polymerized similarly to ethylene, but the polymer, polypropylene, is harder and less flexible than polyethylene due to the presence of its methyl side chains.

$$n H_2C{=}\underset{\displaystyle CH_3}{\underset{|}{CH}} \longrightarrow \text{-}\left(CH_2{-}\underset{\displaystyle CH_3}{\underset{|}{CH}}\right)\text{-}_n$$

Another related material is the polymer of styrene, $C_6H_5CH{=}CH_2$:



Polystyrene is used in packaging, food and household containers (Figure 16.8), combs, and coatings for television lead-in wire. It is rigid, clear, and nontoxic, and it won't dissolve in most aqueous solutions. It can be prepared in the form of expandable beads, which can be molded and shaped. The result is Styrofoam.



FIGURE 16.8 Polystyrene is used in products ranging from clear plastic to Styrofoam.

EXAMPLE 16.9 | Reactions of Alkenes

Write an equation to describe the reaction between 2-hexene and H_2.

Solution:

This is an example of an addition reaction, also called a hydrogenation reaction when the molecule being added is hydrogen. The hydrogen molecule is split apart with the aid of a catalyst, and the hydrogen atoms are added to the carbon atoms involved in the double bond:

$$CH_3CH{=}CHCH_2CH_2CH_3 + H_2 \xrightarrow{Pt} CH_3CH_2CH_2CH_2CH_2CH_3$$

Practice Problem 16.9

Write an equation to describe the reaction that would occur between 2-butene and HBr.

Further Practice: Questions 16.47 and 16.48 at the end of the chapter

Reactions of alkynes are similar to the reactions of alkenes just discussed. Both also undergo combustion reactions. An example is the combustion of acetylene, C_2H_2, when it's used as a fuel in a welding torch:

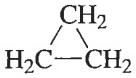
$$2H{-}C{\equiv}C{-}H + 3O_2 \longrightarrow 4CO_2 + 2H_2O$$

16.4 | CYCLIC HYDROCARBONS

Cycloalkanes and Cycloalkenes

Imagine a hydrocarbon chain of single-bonded carbons with a hydrogen atom removed on each end and the two carbon atoms at the end then bonded together. The result would be the ring structure of a saturated cyclic hydrocarbon, or **cycloalkane**, with the general formula C_nH_{2n}. The simplest cycloalkane is cyclopropane, C_3H_6. It is a colorless, flammable gas that is explosive in air. It was once used as a general anesthetic.

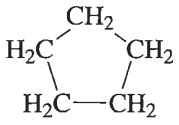




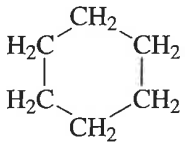
cyclopropane

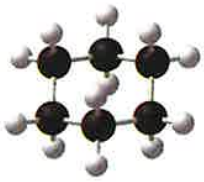
Many cyclic molecules, especially those in which the rings are small, are strained because they cannot adopt the normal tetrahedral bond angles (109.5°) around carbon. Thus, they tend to be somewhat more reactive than the alkanes. The most common cycloalkanes are the low-boiling liquids cyclopentane, C_5H_{10}, and cyclohexane, C_6H_{12}:





cyclopentane



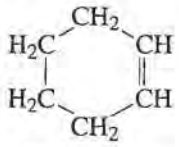


cyclohexane

Cyclopentane is a flammable, colorless liquid that smells like gasoline. Cyclohexane has similar properties and is used as a solvent for lacquers and resins and in paint removers.

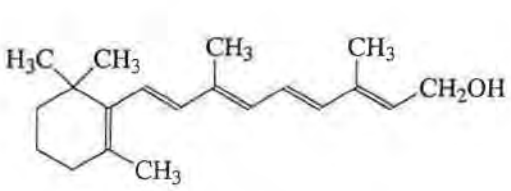
Some alkenes also form ring structures. Molecules containing multiple double bonds and double bonds in ring structures are common in nature. We saw the

example of *cis*- and *trans*-11-retinal that Maya studied; it has five carbon-carbon double bonds, one of which is in a cyclohexene ring:

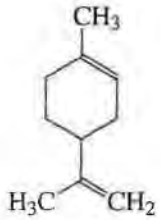


cyclohexene

The official name of retinal has the number 11 in it to designate the position of the double bond that gives rise to the isomers. Some other common examples in nature include vitamin A, required for good vision, and limonene, found in the oils in citrus fruits:



vitamin A



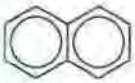
limonene

Aromatic Hydrocarbons

The term *aromatic* was originally assigned to this class of compounds because many have pleasant aromas.

Omar was studying aspirin, which contains a six-carbon ring having alternating single and double bonds. Aisha was interested in amphetamine, which also contains such a six-carbon ring. Although each of these molecules has additional functional groups, both are classified as aromatic. Common aromatic hydrocarbons are cyclic hydrocarbons with alternating single and double carbon-carbon bonds.

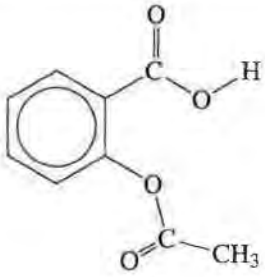
Some aromatic compounds have overlapping multiple six-carbon rings with delocalized double bonds. An example is naphthalene, once widely used as a moth protectant in closets:



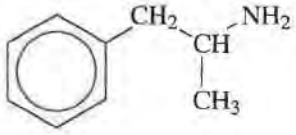
The actual structure of aromatic hydrocarbons such as C_6H_6 has the electrons in the double bonds spread out or delocalized over the entire ring system, as discussed in Section 16.2. Double bonds with electrons delocalized over the entire ring are less reactive than ordinary double bonds. The resulting inertness is referred to as *aromatic character.* The delocalized bonding in benzene, C_6H_6, is usually represented by a circle inside the structure, and the hydrogen atoms, one attached to each carbon atom, are not shown:



Aspirin and amphetamine can be represented by the following structural formulas:

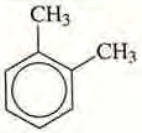
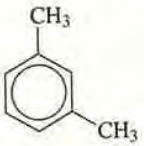
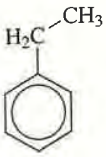
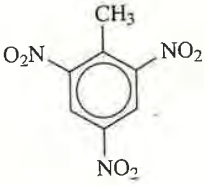


aspirin

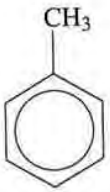


amphetamine

TABLE 16.6 | Some Formulas and Names of Aromatic Hydrocarbons

Structural Formula	Name	Structural Formula	Name
	methylbenzene (or toluene)		1,2-dimethylbenzene
	chlorobenzene		1,3-dimethylbenzene
	aminobenzene (or aniline)		1,4-dimethylbenzene
	ethylbenzene		2,4,6-trinitrotoluene (TNT)
	1-amino-4-chlorobenzene (or 4-chloroaniline)		

Nomenclature of Aromatic Hydrocarbons Derivatives of benzene are generally named with benzene as the base name. Many of these molecules have special names as well. The following molecule, for example, is called both methylbenzene and, more commonly, toluene. A methyl group is substituted for one hydrogen atom on the benzene ring:



Toluene, $C_6H_5CH_3$, is a flammable, colorless liquid that smells like benzene. It is used as a solvent and in the manufacture of benzoic acid, benzaldehyde, dyes, explosives, and numerous organic compounds.

Examine the names of the aromatic compounds in Table 16.6. Can you come up with some rules for naming them?

Did you come up with some rules? How do you name the ring? How do you designate the positions and names of attached groups? If you need help, consult Section 16.9. See if you can apply your rules to naming some aromatic compounds.

EXAMPLE 16.10 | Names of Aromatic Compounds

Name this aromatic compound:

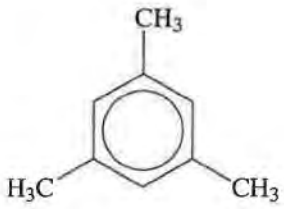


Solution:

The base structure is a six-carbon aromatic ring, so the compound is named as a derivative of benzene. We start numbering the ring at the first attached group (in alphabetical order), ethyl, so the methyl group is at carbon 4. The name is thus 1-ethyl-4-methylbenzene.

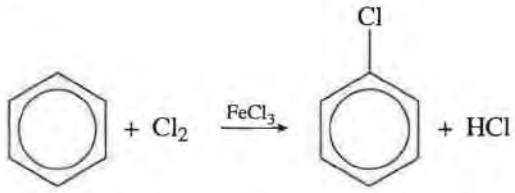
Practice Problem 16.10

Name this aromatic compound:

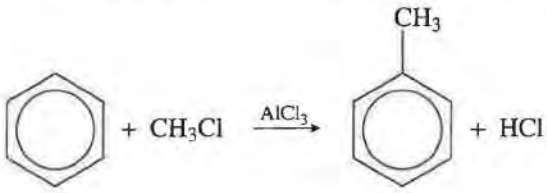


Further Practice: Questions 16.51 and 16.52 at the end of the chapter

Reactions of Aromatic Hydrocarbons Benzene and other aromatic hydrocarbons lack the reactivity of the double bond in alkenes. Normally, benzene does not react at all, but with the right catalyst, it can undergo a *substitution reaction* in which some functional group replaces one or more hydrogens on the benzene ring. For example, benzene reacts with chlorine in the presence of the catalyst $FeCl_3$ to form chlorobenzene:



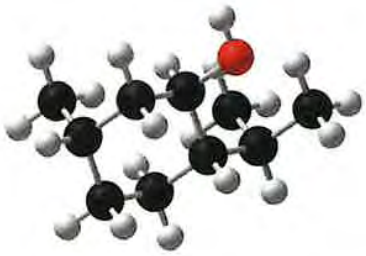
It reacts with chloromethane in the presence of the catalyst $AlCl_3$ to form toluene:

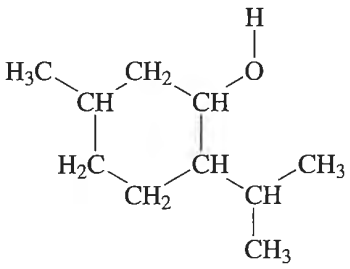


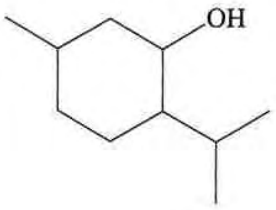
16.5 | ALCOHOLS AND ETHERS

Alcohols

Jennifer's molecule, menthol, is an example of an **alcohol**, which is characterized by the functional group −OH:







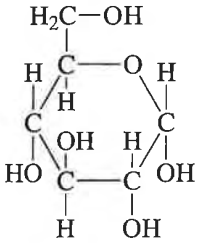
Alcohols are useful substances, and some of them are among the top 50 chemicals produced in the United States. Alcohols produced in large amounts include methanol (CH_3OH), ethylene glycol ($HOCH_2CH_2OH$), isopropanol or 2-propanol (C_3H_7OH), and ethanol (C_2H_5OH).

The −OH group makes alcohols polar, and it can participate in hydrogen-bonding interactions. As a result, even the lower-molar-mass alcohols are liquids, while alkanes of similar size are gases. The short-chain alcohols are water soluble, but solubility decreases as the hydrocarbon portion of the molecule increases in size. For example, C_4H_9OH is only slightly soluble in water, and alcohols containing five or more carbon atoms are essentially insoluble in water. Alcohols of similar size are soluble in one another. For example, methanol dissolves in ethanol. Methanol is a colorless liquid so toxic that breathing its vapors can blind or kill. It is often added to nonbeverage ethanol, to make it unfit for drinking. Modified ethanol can be sold for scientific use and as an additive for gasoline without the federal taxes levied on beverage alcohol. In this role, methanol is called a denaturant, and the modified ethanol is called *denatured alcohol.*

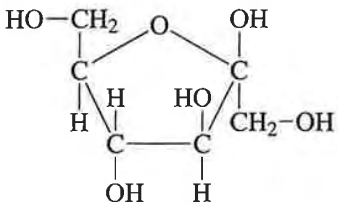
Ethylene glycol, or 1,2-ethanediol, $HO-CH_2-CH_2-OH$, is an alcohol containing two −OH groups. This colorless liquid is commonly used as an antifreeze in automobile radiators.

Because of the polarity and hydrogen bonding of the alcohols, they have significantly higher boiling points than alkanes of comparable mass. For example, the boiling point of CH_3CH_2OH is 78.5°C, but that of $CH_3CH_2CH_3$ is −42.1°C. The alkane molecules are attracted to one another only by London dispersion forces, while alcohol molecules are attracted to one another by stronger dipole forces and hydrogen-bonding forces.

Carbohydrates such as glucose and fructose contain multiple alcohol functional groups:



glucose



fructose

Also notice the presence of an ether functional group, an −O− atom bonded to two carbon atoms in these simple sugars. Larger sugars such as sucrose and maltose contain two glucose and/or fructose units connected by ether linkages. Starch and cellulose are larger polymers of glucose, also connected by ether linkages. Humans

We'll discuss carbohydrates in more detail in Chapter 17.

TABLE 16.7 | Formulas and Names of Some Alcohols

Structural Formula	Name
H_3C-OH	methanol
H_3C-CH_2-OH	ethanol
$H_3C-CH_2-CH_2-OH$	1-propanol
$H_3C-\underset{\displaystyle OH}{\underset{\vert}{CH}}-CH_3$	2-propanol
$H_3C-CH_2-CH_2-CH_2-OH$	1-butanol
$H_3C-CH_2-\underset{\displaystyle OH}{\underset{\vert}{CH}}-CH_3$	2-butanol
$H_3C-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-CH_2-\underset{\displaystyle OH}{\underset{\vert}{CH}}-CH_3$	4-methyl-2-pentanol

and other vertebrate animals have the enzymes necessary to metabolize starch, converting it back to glucose, which cells use as their primary energy source. However, they cannot digest the cellulose present in grass and most other plant foods. Grazers such as cows rely on bacteria that live in their guts to do the job for them.

Nomenclature of Alcohols Examine the names and formulas of some alcohols given in Table 16.7. Try to determine the rules for naming these compounds.

Did you come up with some rules? How do you name the chain? How do you designate the position of any branches? How do you designate the position of the −OH group? If you need help, consult Section 16.9. Try to use your rules to name some alcohols.

EXAMPLE 16.11 | Names of Alcohols

Name this alcohol:

$$H_3C-\underset{\displaystyle OH}{\underset{\vert}{CH}}-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-CH_3$$

Solution:

The longest chain containing the −OH group has four carbons, so this is a derivative of butane. The *–ane* ending is changed to *–anol*, making the name butanol. The −OH group is attached to the second carbon, so the chain name is 2-butanol. A methyl group is attached to carbon 3, so the full name is 3-methyl-2-butanol.

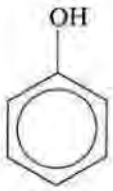
Practice Problem 16.11

Name this alcohol:

$$H_3C-\underset{\displaystyle OH}{\underset{\vert}{CH}}-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-\underset{\displaystyle CH_3}{\underset{\vert}{CH}}-CH_3$$

Further Practice: Questions 16.61 and 16.62 at the end of the chapter

An −OH group may be attached directly to an aromatic hydrocarbon group, as in C_6H_5OH:



phenol

This molecule is known as phenol. The chemistry of the −OH group changes greatly when it is attached directly to an aromatic ring, and such compounds are classified not as alcohols, but as phenols.

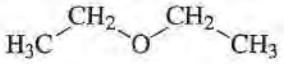
Ethers

An **ether** is an organic molecule that contains an −O− atom bonded to two carbon atoms. Ethers are related to the alcohols, but they have two hydrocarbon groups attached to an oxygen: R−O−R′. The names of the ethers include the names of the attached alkyl groups, listed alphabetically and followed by the word *ether*, as in diethyl ether ($CH_3CH_2-O-CH_2CH_3$). The ethers tend to be more volatile than the corresponding alcohols. Consider the boiling points of an alkane, an ether, and an alcohol of comparable mass:

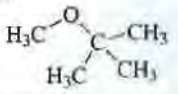
Formula	$CH_3CH_2CH_3$	CH_3OCH_3	CH_3CH_2OH
Boiling point (°C)	−42.1	−23.6	78.5

Ethers are polar molecules, so they have stronger intermolecular forces and higher boiling points than alkanes. With no hydrogen atom attached to the oxygen atom in an ether, there is no hydrogen bonding, so it has a lower boiling point than the comparable alcohol.

One of the molecules first used as a general anesthetic was diethyl ether:



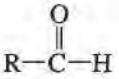
MTBE, or methyl tertiary-butyl ether, is added to some formulations of gasoline to ensure more complete combustion:



16.6 | ALDEHYDES AND KETONES

Aldehydes

Earlier we discussed Maya's favorite molecules, *cis*- and *trans*-retinal, as examples of alkenes. However, these molecules contain the −CHO functional group, so they are also aldehydes. An **aldehyde** contains a C=O (carbonyl) group in which at least one hydrogen atom is attached to the carbon atom. Aldehydes have the general formula R−CHO, in which the hydrogen and the oxygen are both bonded to the carbon, and the oxygen-carbon bond is a double bond:



The most common aldehydes are formaldehyde, HCHO, and acetaldehyde, CH_3CHO. Formaldehyde (common name), or methanal (IUPAC name), is a colorless gas with an irritating odor. It is highly soluble in water and is commonly sold as a

The term *resin* was originally used to describe the thick, odoriferous sap exuded by cuts in the bark of pine trees. It is now used in a more generic sense to include the many synthetic, amorphous solid materials that are used in plastics, varnishes, inks, and medicines.



FIGURE 16.9 Urea-formaldehyde foam used for insulation is formed by mixing two liquids. They react to yield a polymer and a gas. The gas causes the polymer to expand and form a foam. In a short time the foam hardens into a rigid solid containing a lot of trapped gas.



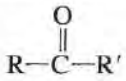
FIGURE 16.10 During the Tollens test for aldehydes, silver(I) ion is reduced to metallic silver and the aldehyde is oxidized. Here the reaction with formaldehyde creates a silver mirror on the flask. Ketones are not easily oxidized and do not react in this way.

37% solution known as *formalin*. Formalin causes the coagulation of proteins, making it useful as a preservative of tissues. Formaldehyde is also used with phenol to form polymer resins, which are used in plastics and as adhesives in plywood and particleboard. Urea-formaldehyde foam (Figure 16.9) was once widely used for house insulation, until it was banned because of the toxic properties of formaldehyde.

Acetaldehyde (common name), or ethanal (IUPAC name), CH_3CHO, is a volatile, colorless liquid that is soluble in water. It has a characteristic, pungent odor much like that of cut green apples. It is toxic and flammable. Acetaldehyde is used in perfumes, aniline dyes, plastics, and synthetic rubber. Aldehydes are easily oxidized, so they make good reducing agents. The Tollens test uses this property to indicate the presence of aldehydes (Figure 16.10).

Ketones

Ketones resemble aldehydes in that both contain a carbon double-bonded to an oxygen, or C=O. In an aldehyde, this group is located on an end carbon. In a **ketone,** however, the organic groups, often alkyl groups—such as methyl or ethyl here designated R and R′—lie on either side of the C=O group:

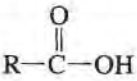


The simplest and most important of the ketones is acetone, $(H_3C)_2C{=}O$, which is officially named propanone. Acetone is a colorless, highly flammable liquid with a characteristic pungent odor. It is readily soluble in water. Acetone is widely used as a solvent for fats, oils, waxes, resins, rubber, plastics, lacquers, varnishes, and rubber cements. It is the principal ingredient in some fingernail polish removers.

16.7 | CARBOXYLIC ACIDS AND ESTERS

Carboxylic Acids

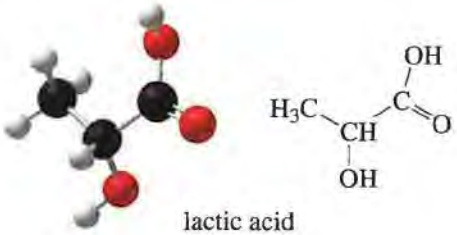
Omar's favorite molecule, aspirin, contains a carboxylic acid functional group. The characteristic functional group in a **carboxylic acid** molecule is a carbon atom double-bonded to an oxygen atom and also bonded to an −OH group, $-CO_2H$:



Carboxylic acids are typically weak acids, and their reactions often involve loss of H^+.

The simplest carboxylic acid is formic acid (methanoic acid), HCO_2H, which was first obtained from red ants. It is the chemical agent that causes ant bites, bee stings, and nettle stings to be so painful. Formic acid is a colorless, water-soluble liquid with a pungent odor. It is used in dyeing wool, dehairing and tanning hides, electroplating, coagulating rubber latex, and regenerating old rubber.

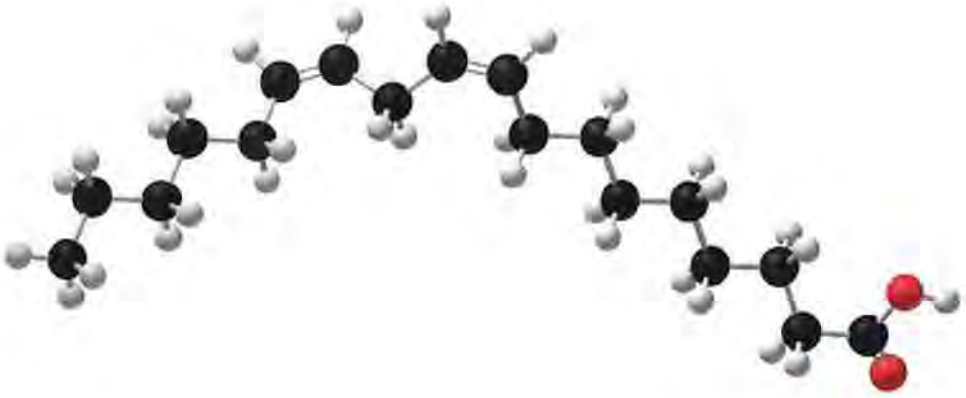
Another example is the lactic acid produced in muscles from the metabolism of glucose:

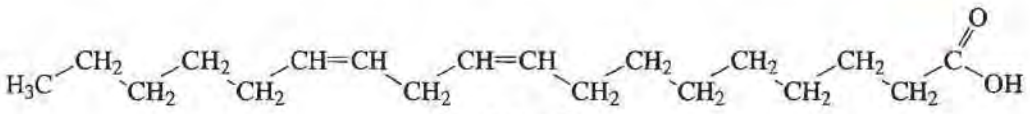


lactic acid

It causes muscle soreness when its concentration builds up. The lactic acid molecule also contains an alcohol functional group.

Another carboxylic acid is linoleic acid. It is a nutritionally essential fatty acid that is also used in manufacturing paints. It also contains two alkene functional groups:





linoleic acid

The most common carboxylic acid is acetic acid, CH_3CO_2H, which is produced as a solution in water, called vinegar, when cider ferments. Fermentation causes the sugar to form ethanol in the presence of catalytic yeast enzymes. In the presence of air, the ethanol oxidizes further to acetic acid:

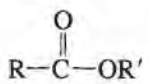
$$CH_3CH_2OH + O_2 \xrightarrow{\text{yeast}} CH_3CO_2H + H_2O$$

Pure acetic acid is a liquid that has a pungent odor and burns the skin. It is completely soluble in water and is an excellent solvent for many organic compounds, including resins and oils. Acetic acid is used in the synthesis of rubber and various acetate plastics. It is used in printing calico and dyeing silk.

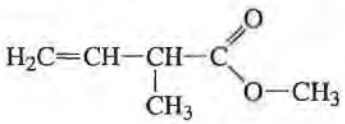
Many breathalyzer test devices use the oxidizing agent $K_2Cr_2O_7$ (an orange compound) to oxidize ethanol to acetic acid. In the presence of ethanol, $K_2Cr_2O_7$ is reduced to a green compound. The degree of color change indicates the concentration of ethanol in a person's breath.

Esters

An **ester** is a compound with the general formula RCO_2R'. In esters, one of the organic groups (often alkyl) is directly attached to the carbon, another is bonded to one of the oxygens, and the other oxygen is double-bonded to the carbon:



The aspirin that interests Omar is both a carboxylic acid and an ester. Another example of an ester is methyl methacrylate, which is used to form the polymer Plexiglas:



A condensation reaction involves the combination of two molecules accompanied by the elimination of H_2O.

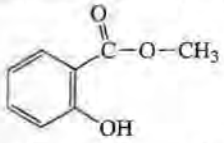
Esters can be formed by a *condensation reaction* of a carboxylic acid with an alcohol. Water is removed in the process, as shown here for the formation of ethyl acetate by the reaction of acetic acid with ethanol:



These reactions are reversible and usually require an acid catalyst. Addition of a base neutralizes and removes the carboxylic acid, causing the equilibrium reaction to shift to the left. Most esters are volatile liquids with pleasant odors and flavors. The natural fragrances of flowers and plants often come from esters. Esters are often used as flavoring agents in synthetic and processed foods.

Many fruits and flowers contain esters or mixtures of esters that give them their characteristic odors. Some examples are

methyl salicylate wintergreen

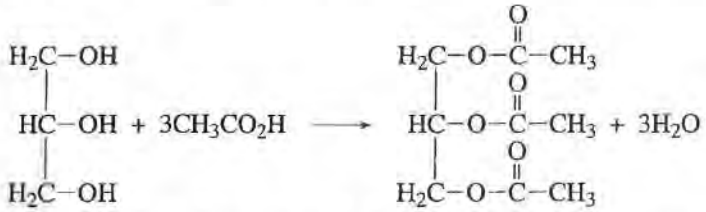


isoamyl acetate banana
$(CH_3)_2CHCH_2CH_2O_2CCH_3$

octyl acetate orange
$CH_3(CH_2)_6CH_2O_2CCH_3$

methyl butyrate apple
$CH_3O_2CCH_2CH_2CH_3$

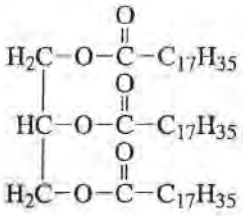
The common animal and vegetable oils and fats are esters, mostly based on an alcohol called glycerol (or 1,2,3-propanetriol). Glycerol reacts with carboxylic acids to form an ester called a *triglyceride:*



Esters that form from the reaction of glycerol with high-molar-mass acids are liquid oils or solid fats, depending on the number of double bonds in the alkyl group of the acid. (Such carboxylic acids are called *fatty acids.*) If the fatty acid is saturated (no double bonds), the triglyceride is a solid, or *fat;* and if the fatty acid is highly unsaturated, with many carbon-carbon double bonds, the triglyceride is a liquid, or *oil.* The greater the unsaturation, the more liquidlike the triglyceride. Common fats include lard, tallow, and milk fat. Common oils are linseed, cottonseed, coconut, soybean, peanut, corn, and olive oils.

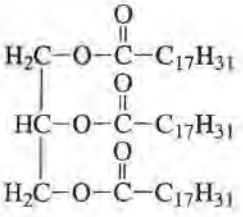
Fats and oils are used in the body to synthesize essential tissue constituents. The body also uses them as energy sources. Because they are insoluble, they can be deposited in cells and can remain there unchanged until energy is needed.

A solid ester forms from glycerol and stearic acid, $C_{17}H_{35}CO_2H$, which is a saturated fatty acid:

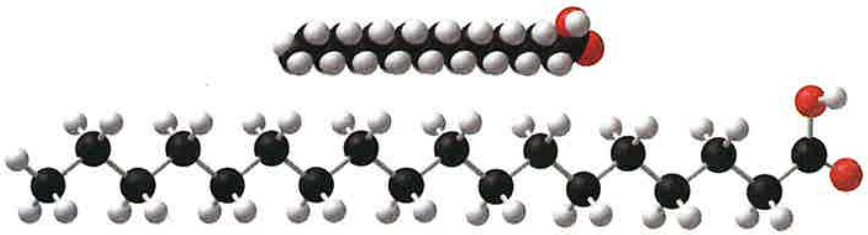


Triglycerides are the principal form of fats circulating in the human bloodstream. Most body fat also is in the form of triglycerides. A simple blood test measures the amount of triglycerides in blood. It is one measure of the risk of developing coronary artery disease.

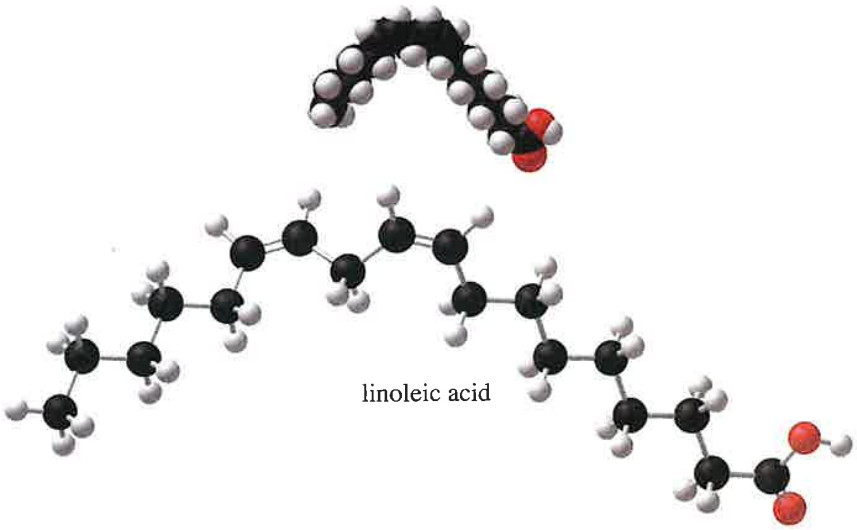
A liquid ester forms from glycerol and linoleic acid, $C_{17}H_{31}CO_2H$, which is an unsaturated fatty acid:



Why do these two esters—one a solid and the other a liquid at room temperature—have different melting points? The answer lies in the structure of the carboxylic acids from which the esters were derived:



stearic acid



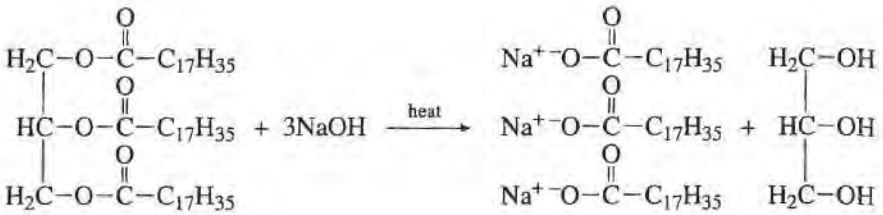
linoleic acid

Linoleic acid is bent because of a cis arrangement around a carbon-carbon double bond. Which of these acids would have the greater intermolecular forces? Recall the arguments we made for greater intermolecular forces in straight-chain hydrocarbons than in branched-chain hydrocarbons. Since the straight-chain molecules can interact to a greater extent than ones with bent shapes, they have greater intermolecular forces and, as a result, higher melting points. The same logic applies here. Linoleic acid has a melting point of –12°C, while stearic acid has a melting point of 70°C. Thus, the esters derived from them have melting points that differ in the same way.

The glyceride of linoleic acid is found in the oils of many seeds and vegetables. Stearic acid comes from animal fats and oils. Do all solid fats come from animals, while liquid fats come from vegetables? Generally, yes, but oils from vegetables can be converted in the laboratory into solids by the process of hydrogenation, in which hydrogen molecules add across the double bonds in the presence of an appropriate catalyst. This is how margarine is made.

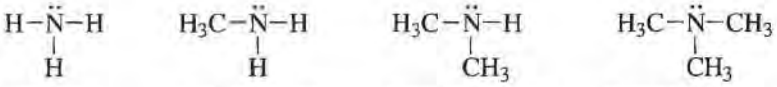
Hydrogenation: $\underset{\text{unsaturated (oil)}}{C_3H_5(C_{17}H_{31}CO_2)_3} + 6H_2 \longrightarrow \underset{\text{saturated (solid)}}{C_3H_5(C_{17}H_{35}CO_2)_3}$

The sodium compounds of the carboxylic acids obtained from fats and oils are used in soaps. Their anions have polar ends that attract water and nonpolar ends that attract grease and oil, dispersing them into the water. Soap is prepared from animal fats by *saponification*, a reaction with sodium hydroxide (lye) that cleaves the ester linkages and converts the ester to sodium carboxylates and glycerol:



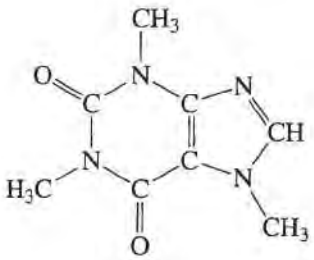
16.8 | AMINES

Aisha was interested in the aromatic molecule amphetamine. Because it contains a nitrogen atom singly bonded to a carbon, amphetamine is also an example of an organic amine. Another aromatic amine is aniline, $C_6H_5NH_2$, which is used in dyes. An **amine** can be considered a derivative of ammonia, NH_3, with one or more hydrogen atoms replaced by organic groups. Examples of some simple amines are shown here.

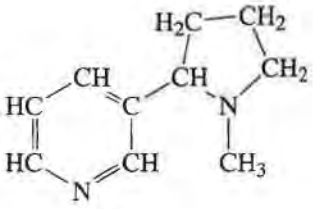


ammonia methylamine dimethylamine trimethylamine

Many amines are classified as alkaloids, which are basic nitrogen-containing compounds that come from plants. Many alkaloids have physiological effects on humans, most of which are toxic. Some examples of alkaloids are caffeine and nicotine, which interact with living cells and affect their functioning.



caffeine



nicotine

It is also possible to make quaternary ammonium compounds such as tetramethyl ammonium chloride, $(CH_3)_4N^+Cl^-$. This can be considered a derivative of $NH_4^+Cl^-$, with all four hydrogen atoms replaced by methyl groups.

Like ammonia, most amines act as bases due to the presence of an unshared electron pair on the nitrogen atom. Most amines are weak bases in water and react completely with strong acids, such as HCl, to form ammonium salts:

$$R{-}NH_2 + HCl(aq) \longrightarrow R{-}NH_3^+Cl^-(aq)$$

16.9 | ORGANIC NOMENCLATURE

Organic nomenclature helps us communicate structures without drawing a complete structural formula. Although many names are in common use, the IUPAC system is the standard system used worldwide. Once you learn the basics of the IUPAC system, you can easily name a variety of organic compounds with various functional groups.

Alkanes

To use IUPAC nomenclature, start with the names of the alkyl groups derived from the alkanes by removal of one hydrogen atom, with a general formula C_nH_{2n+1}. The names of the alkyl groups are similar to those of the alkanes given in Table 16.2, except that the *–ane* ending is replaced by *–yl*. Thus, $-CH_3$ is methyl and $-C_2H_5$ is ethyl.

A few rules allow us to achieve consistent nomenclature for the alkanes.

1. Determine the *longest* continuous carbon chain in the hydrocarbon. The name of this alkane constitutes the parent name.
2. Number the carbon atoms in the parent hydrocarbon from the end nearer a branch, so that branch will have the lowest possible position number.
3. Branches attached to the parent chain are named as alkyl groups.
4. The name consists of the position number, a hyphen, the branch name, and the straight-chain name. If there is more than one branch, name them in alphabetical order. If the branches are the same alkyl group, precede the name with a prefix (*di–*, *tri–*, etc.) indicating the total number of that alkyl group in the molecule, and separate the position numbers with commas.

The cycloalkanes are named in the same way as the straight-chain alkanes, but with the prefix *cyclo–*. Thus, C_6H_{12}, in which six carbons are bonded in a ring, is named cyclohexane.

Alkenes and Alkynes

The names of the alkenes use the alkane root, with the ending *–ene* replacing the *–ane* ending. If a double bond has more than one possible position, a number indicates its location. The carbon atoms are numbered from the end nearer the double bond. For the following compound, the carbon chain is numbered from left to right, which places the C=C bond between carbon atoms 1 and 2:

$$\overset{1}{H_2C}=\overset{2}{CH}-\overset{3}{CH_2}-\overset{4}{CH_3}$$

This molecule has four carbon atoms, so it is derived from butane. It has a double bond, so the *–ane* ending is replaced by *–ene,* giving it the name butene. The position of the double bond is designated by the lowest-numbered carbon involved in that bond, so this molecule is 1-butene. Similarly the molecule $CH_3-CH=CH-CH_3$ is named 2-butene.

The names for side chains on an alkene follow the rules for naming alkyl groups. Side chain names have the *–yl* ending, and a number designates the position of attachment. Again, numbering starts at the end nearest the double bond.

A molecule may have more than one double bond. In such a case, an appropriate prefix (such as *di–*) is placed just before the *–ene* ending to indicate the number of double bonds. For example, $H_2C=CH-CH=CH_2$ is named 1,3-butadiene.

Alkynes are named the same as alkenes, but using a *–yne* ending instead of an *–ene* ending.

If the double bond is closest to the right side of the molecule, as written, then we start counting from that end of the molecule:

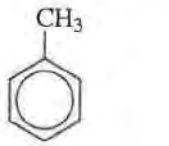
$$\overset{4}{H_3C}-\overset{3}{CH_2}-\overset{2}{CH}=\overset{1}{CH_2}$$

Some alkenes, such as $CH_3-CH=CH-CH_3$, have geometric isomers. This molecule would be named *cis*-2-butene or *trans*-2-butene.

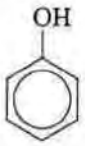
Aromatic Hydrocarbons

An aromatic group formed by removal of a hydrogen atom from a benzene ring is called an **aryl group**, which is the aromatic analog of an alkyl group. The aryl group formed from benzene, C_6H_5, is called phenyl. For example, $C_6H_5C{\equiv}CH$ is commonly known as phenylacetylene.

If two or more positions on a benzene ring have substituted groups, the positions are indicated by numbers, counted from the location of the first substituted group. If two methyl groups are substituted on benzene, the molecule is dimethylbenzene or xylene. There are three isomers: 1,2-dimethylbenzene, 1,3-dimethylbenzene, and 1,4-dimethylbenzene. These and several other examples are summarized here:



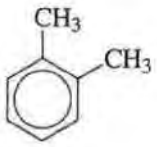
methylbenzene
(toluene)



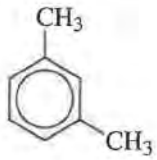
hydroxybenzene
(phenol)



aminobenzene
(aniline)



1,2-dimethylbenzene



1,3-dimethylbenzene



1,4-dimethylbenzene

Other Naming Conventions

The other major classes of organic compounds are named by similar rules, which are summarized in Table 16.8.

TABLE 16.8 | Rules for Naming Other Organic Molecules

Class of Compound	Rules for Naming
alcohols	The common names of the alcohols consist of the alkyl group name followed by the word *alcohol*. The IUPAC names use the alkane root and replace the *–e* ending with *–ol*. Thus, CH_3OH is methyl alcohol or methanol, C_2H_5OH is ethyl alcohol or ethanol, and C_3H_7OH is propyl alcohol or propanol.
ethers	The common names of the ethers consist of the names of the substituted alkyl groups, listed alphabetically, and followed by the word *ether*, as in diethyl ether, $C_2H_5OC_2H_5$, or ethylmethyl ether, $C_2H_5OCH_3$.
aldehydes	Common names use the ending *–aldehyde*, such as in formaldehyde, HCHO, or acetaldehyde, CH_3CHO. IUPAC names use the alkane name with the ending *–al*, such as in methanal, HCHO, or ethanal, CH_3CHO.
ketones	Common names use the names of the two alkyl groups attached to the C=O group, followed by the word *ketone*. IUPAC names use the alkane name with the ending *–one*. The solvent commonly called acetone, $(CH_3)_2C{=}O$, is dimethyl ketone or propanone.
carboxylic acids	Carboxylic acids often have common names, such as formic acid, HCO_2H, and acetic acid, CH_3CO_2H. The IUPAC names use the alkane name with the ending *–oic acid*. The two examples are methanoic acid and ethanoic acid, respectively.
esters	Esters have an alkyl group replacing the acidic hydrogen in a carboxylic acid. They are named in the same way as the carboxylic acids from which they are derived, using an *–oate* ending. Thus, $HCO_2CH_2CH_3$ is ethyl methanoate (or ethyl formate), and $CH_3CO_2CH_3$ is methyl ethanoate (or methyl acetate).
amines	Amines are commonly named by listing the alkyl groups attached to nitrogen, followed by the word *amine*. For example, $NH_2CH_2CH_3$ is ethylamine and $NH(CH_3)_2$ is dimethylamine.

SUMMARY

Organic compounds contain carbon, hydrogen, and sometimes include other elements. Carbon makes up great numbers of compounds because its atoms form four strong bonds. Carbon can bond to other carbon atoms to form chains or rings. It can have single, double, or triple bonds, and it bonds easily to oxygen and nitrogen. Many organic compounds exist as isomers with the same molecular formula but different structures, physical properties, and reactivities. For this reason, we usually use structural formulas, condensed structural formulas, or line structures when describing organic compounds.

Organic compounds are classified by their shape, the kinds of bonds they contain, and the presence of functional groups. Functional groups impart specific properties to organic compounds; they are generally the part of the molecule that undergoes characteristic reactions.

Hydrocarbons contain only carbon and hydrogen, but can have single, double, or triple bonds between carbon atoms. Hydrocarbons are generally obtained from petroleum. The simplest hydrocarbons are the alkanes, which have only single bonds between carbon atoms. Alkanes are sometimes called saturated hydrocarbons because they have the maximum possible number of hydrogen atoms (C_nH_{2n+2} for chain alkanes). Because of the lack of functional groups, alkanes are fairly unreactive, with the exception of combustion reactions and other reactions that must be initiated by heat or light.

Alkenes contain double bonds and alkynes contain triple bonds. Alkenes and alkynes undergo reactions, such as polymerization and addition reactions, which result in the loss of multiple bonds. The most common aromatic hydrocarbons contain a benzene ring, C_6H_6. The delocalization of electrons in benzene makes it less reactive than alkenes.

Alcohols (R−OH) and ethers (R−O−R′) contain carbon-oxygen single bonds. Alcohols are liquids or solids at room temperature because of hydrogen bonding. Reactions of alcohols and ethers occur at the oxygen part of the molecule. Alcohols can be oxidized to aldehydes, and then further to carboxylic acids or to ketones.

Aldehydes and ketones contain a carbonyl group (C=O). In aldehydes the carbonyl group has at least one hydrogen atom bonded to the carbon atom of the carbonyl group, so it is found at the end of carbon chains.

Carboxylic acids (RCO_2H) are the "organic acids," because their reactivity often involves the loss of H^+. Esters (RCO_2R') form from carboxylic acids and alcohols. Esters are responsible for many of the pleasant odors of plants.

Amines contain one to three carbon atoms with single bonds to a nitrogen atom (RNH_2, R_2NH, or R_3N). Amines generally act as bases in reactions and are found in drugs such as caffeine and nicotine.

Organic nomenclature helps us communicate without drawing structures. Although many common names are used frequently, the IUPAC system is the international standard.

KEY TERMS

acyclic hydrocarbon (16.2)
alcohol (16.5)
aldehyde (16.6)
aliphatic hydrocarbon (16.2)
alkane (16.3)
alkene (16.3)
alkyl group (16.3)
alkyne (16.3)
amine (16.8)
aromatic hydrocarbon (16.2)
aryl group (16.9)
carboxylic acid (16.7)
cyclic hydrocarbon (16.2)
cycloalkane (16.4)
ester (16.7)
ether (16.5)
functional group (16.1)
hydrocarbon (16.1)
isomer (16.1)
ketone (16.6)
organic chemistry (16.1)
saturated hydrocarbon (16.2)
unsaturated hydrocarbon (16.2)

QUESTIONS AND PROBLEMS

The following questions and problems, except for those in the *Additional Questions* section, are paired. Questions in a pair focus on the same concept. Answers to the odd-numbered questions and problems are in Appendix D.

Matching Definitions with Key Terms

16.1 Match the key terms with the descriptions provided.
(a) a compound having the same composition and formula as another compound, but a different structure
(b) an organic compound containing the functional group $-OH$
(c) a condensation product of carboxylic acids and alcohols, having the general formula RCO_2R'
(d) an organic compound containing the functional group $-CHO$
(e) a hydrocarbon that has a ring or closed-chain structure
(f) a hydrocarbon that contains only carbon-carbon single bonds
(g) a group derived from an alkane by removal of one hydrogen atom
(h) an alkane that contains a ring structure
(i) a hydrocarbon that contains localized single, double, or triple bonds
(j) a small molecule that contains both the $-NH_2$ and the $-CO_2H$ functional groups
(k) the study of organic compounds
(l) a hydrocarbon that does not contain all the hydrogen atoms that could possibly be included

16.2 Match the key terms with the descriptions provided.
(a) a hydrocarbon that contains one or more localized carbon-carbon double bonds
(b) an organic compound containing an oxygen double-bonded to a carbon, which is also bonded to two other carbons: $R_2C{=}O$
(c) an organic compound containing the functional group $-CO_2H$
(d) a reactive part of a molecule that undergoes characteristic reactions
(e) a hydrocarbon containing one or more triple bonds
(f) an organic compound that contains the $-O-$ functional group attached to two organic groups: $R{-}O{-}R'$
(g) an aromatic group formed by removal of a hydrogen from a cyclic aromatic hydrocarbon
(h) a cyclic hydrocarbon with delocalized bonding, which imparts special stability
(i) a compound that contains only carbon and hydrogen
(j) a hydrocarbon that contains all the hydrogens it can hold
(k) a hydrocarbon that has a straight- or branched-chain structure with only single bonds
(l) an organic base with one or more alkyl groups bonded to a nitrogen atom

Representations of Organic Molecules

16.3 Describe the type of bonds in each of the following molecules.
(a) CH_3CH_3 (b) CH_2CH_2 (c) CHCH

16.4 Describe the type of bonds in each of the following molecules.
(a) $CH_3CH_2CH_3$ (b) CH_3CHCH_2 (c) $CHCCH_3$

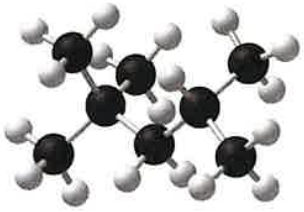
16.5 Identify and classify the functional group in each of the following molecules.
(a) CH_3NH_2
(b) $H_3C-\overset{\overset{O}{\|}}{C}-CH_3$
(c) CH_3CH_2OH
(d) $CH_2{=}CHCH_3$
(e) $CH_3CH_2CH_2CO_2H$
(f) $CH_3C{\equiv}CH$
(g) $H_3C-CH_2-\overset{\overset{O}{\|}}{C}-H$
(h) $CH_3CH_2OCH_3$
(i) $H_3C-CH_2-CH_2-\overset{\overset{O}{\|}}{C}-O-CH_3$

16.6 Identify and classify the functional group in each of the following molecules.
(a) $CH_3CH_2C{\equiv}CH$
(b) $H_3C-\overset{\overset{O}{\|}}{C}-O-CH_3$
(c) $CH_3CH_2NH_2$
(d) $H_3C-CH_2-\overset{\overset{O}{\|}}{C}-CH_3$
(e) $CH_3CH_2OCH_2CH_3$
(f) $H_3C-\overset{\overset{OH}{|}}{C}H-CH_3$
(g) CH_3CO_2H
(h) $CH_3CH_2CH{=}CH_2$
(i) $H_3C-\overset{\overset{O}{\|}}{C}-H$

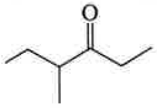
16.7 From the ball-and-stick model, determine the molecular formula and draw the structural formula and the line structure for this molecule.

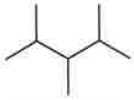


16.8 From the ball-and-stick model, determine the molecular formula and draw the structural formula and the line structure for this molecule.



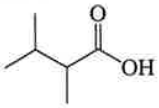
16.9 Write molecular formulas for each of the following.

(a) 

(b) 

16.10 Write molecular formulas for each of the following.

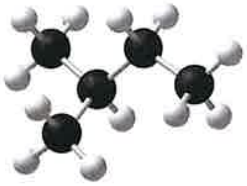
(a) 

(b) 

16.11 Draw line structures for each of the following.
(a) $CH_3(CH_2)_6CH_3$
(b) $CH_3CH_2COCH_2CH_3$
(c) $(CH_3)_2CHCH_2CH_2CHCH_2$
(d) $CH_3(CH_2)_5CHO$
(e) $(CH_3)_4C$

16.12 Draw line structures for each of the following.
(a) $CH_3(CH_2)_8CH_3$
(b) $CH_3CH_2CH_2COCH_2CH_3$
(c) $CH_3CH_2CHCHCH(CH_3)_2$
(d) $CH_3(CH_2)_5CO_2H$
(e) $(CH_3)_3CH$

16.13 Write structural formulas and draw the line structures for each of the following.

(a) 

(b) 

(c) 

16.14 Write structural formulas and draw the line structures for each of the following.

(a) 

(b) 

(c) 

16.15 Draw a perspective drawing for the following molecules.
(a) CH_4
(b) CH_3CH_3
(c) CH_2ClCH_2Cl

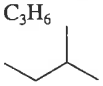
16.16 Draw a perspective drawing for the following molecules.
(a) CH_3Cl
(b) CCl_3CH_3
(c) CH_3CHCl_2

Hydrocarbons

16.17 Distinguish between alkanes, alkenes, and alkynes.

16.18 Write line structures for an alkane, an alkene, and an alkyne, each of which contains four carbon atoms.

16.19 Label each of the following molecules as saturated or unsaturated.
(a) 
(b) C_4H_6
(c) $CH_3CHCHCH_2CH_3$

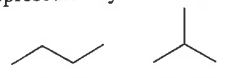
16.20 Label each of the following molecules as saturated or unsaturated.
(a) $CH_3CH_2CHCH_2$
(b) C_3H_6
(c) 

16.21 What distinguishes an aromatic hydrocarbon from other hydrocarbons?

16.22 What feature is common to all aromatic hydrocarbons?

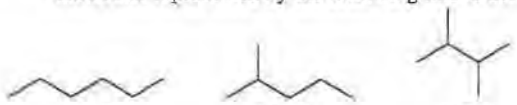
Acyclic Hydrocarbons

16.23 Explain the difference in boiling points of the two molecules represented by the following line structures.



–0.5°C –11.6°C

16.24 Explain the difference in boiling points of the three molecules represented by the following line structures.



69°C 60.3°C 58°C

16.25 Which compound in each group is expected to have the strongest intermolecular forces?
(a) CH_4, CH_3CH_3, $CH_3CH_2CH_3$
(b) $CH_3(CH_2)_4CH_3$, $CH_3CH(CH_3)CH(CH_3)CH_3$, $CH_3CH_2C(CH_3)_2CH_3$

16.26 Which compound in each group is expected to have the strongest intermolecular forces?
(a) CH_2CH_2, CH_3CHCH_2, $CH_3CHCHCH_3$
(b) $CH_3CH_2CH_2CH_2CH_3$, $CH_3C(CH_3)_3$, $CH_3CH_2CH(CH_3)_2$

16.27 What is the name of each of the following straight-chain alkanes?
(a) CH_3CH_3
(b) $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
(c) $CH_3CH_2CH_3$
(d) $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

16.28 What is the name of each of the following straight-chain alkanes?
(a) $CH_3CH_2CH_2CH_3$
(b) $CH_3CH_2CH_2CH_2CH_3$
(c) $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
(d) $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
(e) $CH_3CH_2CH_2CH_2CH_2CH_3$

16.29 Draw all of the possible isomers with the formula C_6H_{14}.

16.30 Draw all of the possible isomers with the formula C_5H_{12}.

16.31 Write IUPAC names for the following compounds.
(a) $CH_3CH_2CH_2CH(CH_3)CH_3$
(b) $CH_3CH(CH_3)CH_2CH_3$
(c) $CH_3CH_2CH(CH_3)CH_3$

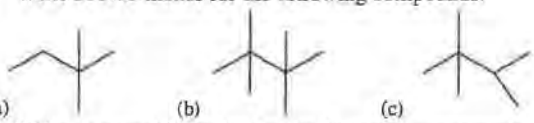
16.32 Write IUPAC names for the following compounds.
(a) $CH_3CH(CH_3)_2$
(b) $CH_3CH_2CH(CH_3)CH_2CH_2CH_3$
(c) $CH_3CH_2CH(CH_2CH_3)CH_2CH_2CH_3$

16.33 Write IUPAC names for the following compounds.



(a) (b) (c)

16.34 Write IUPAC names for the following compounds.



(a) (b) (c)

16.35 Write structural formulas for the following compounds.
(a) 2-methylpropane
(b) 3,3-diethyl-2-methylpentane
(c) 2,4-dimethyloctane
(d) 3-ethyl-3-methylhexane

16.36 Write structural formulas for the following compounds.
(a) 2-methylbutane
(b) 3,3-diethyl-2-methylhexane
(c) 3,5-dimethylnonane
(d) 3-ethyl-2,2-dimethylhexane

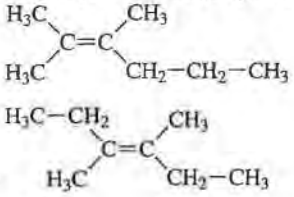
16.37 Predict the products of the following reactions.
(a) $CH_3CH_2CH_3 + Cl_2 \xrightarrow{\text{heat}}$
(b) $(CH_3)_3CCH_2CH_2CH_3 + O_2 \xrightarrow{\text{heat}}$
(c) $CH_3CH_3 + Br_2 \xrightarrow{\text{light}}$

16.38 Predict the products of the following reactions.
(a) $CH_4 + Cl_2 \xrightarrow{\text{heat}}$
(b) $CH_3CH_2CH_3 + O_2 \xrightarrow{\text{heat}}$
(c) $CH_3CH_2CH_3 + Br_2 \xrightarrow{\text{light}}$

16.39 What is an alkene?

16.40 What is an alkyne?

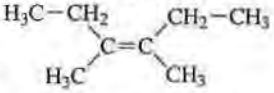
16.41 Write IUPAC names for the following compounds.
(a) $CH_3CH_2CH{=}CHCH(CH_3)CH_2CH_3$



(b)

(c)

16.42 Write IUPAC names for the following compounds.
(a) $(CH_3)_2C{=}CHCH_3$
(b) $CH_3CH{=}CHC(CH_3)_3$



(c)

16.43 Which of the following compounds can exist as cis and trans isomers?
(a) $CH_2{=}CH_2$
(b) $(CH_3)_2C{=}CHCH_3$
(c) $CH_3CH_2C(CH_3){=}C(CH_3)CH_2CH_3$

16.44 Which of the following compounds can exist as cis and trans isomers?
(a) $CH_3CH{=}CHCH_3$
(b) $CH_3CH{=}CH_2$
(c) $(CH_3)_2C{=}C(CH_3)_2$

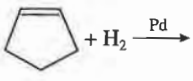
16.45 Write line structures for the following compounds.
(a) 3-methyl-1-butene
(b) propyne
(c) *trans*-2-butene
(d) *cis*-3-hexene

16.46 Write line structures for the following compounds.
(a) 2,3,3-trimethyl-1-hexene
(b) 1-butyne
(c) *cis*-2-butene
(d) *trans*-3-hexene

16.47 Complete the following reactions.

(a) $CH_3CH{=}CHCH_3 + HBr \longrightarrow$

(b) $CH_3CH_2CH{=}CHCH_3 + Br_2 \longrightarrow$

(c)  $+ H_2 \xrightarrow{Pd}$

16.48 Complete the following reactions.

(a) $CH_3CH_2CH{=}CHCH_2CH_3 + HBr \longrightarrow$

(b) $CH_2{=}CHCH_2CH_3 + Br_2 \longrightarrow$

(c)  $+ H_2 \xrightarrow{Pd}$

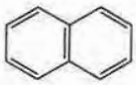
Cyclic Hydrocarbons

16.49 What feature is common to all aromatic hydrocarbons?

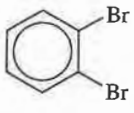
16.50 Which of the following are aromatic?

(a) $CH_2{=}CHCH{=}CH_2$

(b) 

(c) 

16.51 Name the following compounds.

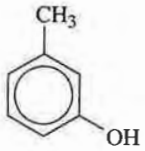


(a)

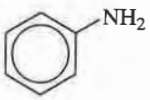


(b)

16.52 Name the following compounds.



(a)



(b)

16.53 Write structural formulas for the following compounds.

(a) 1,2,4-trimethylbenzene

(b) chlorobenzene

16.54 Write structural formulas for the following compounds.

(a) 2,5-dichloro-1-ethylbenzene

(b) 2-ethyl-1,3-dimethylbenzene

16.55 Write an equation for the reaction of benzene, C_6H_6, with chlorine, Cl_2, in the presence of an $FeCl_3$ catalyst.

16.56 Write an equation for the reaction of benzene, C_6H_6, with chloromethane, CH_3Cl, in the presence of an $AlCl_3$ catalyst.

Alcohols and Ethers

16.57 Which of the following compounds are alcohols?

(a) CH_3OCH_3

(b) CH_3CH_2OH

(c) $CH_3CH(OH)CH_3$

(d) CH_3COCH_3

(e) CH_3CH_2CHO

16.58 Which of the following compounds are alcohols?

(a) CH_3CHO

(b) $(CH_3)_2C(OH)CH_3$

(c) $CH_3CH_2COCH_3$

(d) $CH_3CH_2OCH_2CH_3$

(e) $CH_3CH_2CH_2CH_2OH$

16.59 An alcohol has the formula C_4H_9OH. How many isomers can exist for this alcohol?

16.60 An alcohol has the formula $C_5H_{11}OH$. How many isomers can exist for this alcohol?

16.61 Name the following compounds using the IUPAC system.

(a) $(CH_3)_2CHCH_2OH$

(b) $CH_3CH(OH)CH_2CH_3$

16.62 Name the following alcohols using the IUPAC system.

(a) $CH_3CH_2CH(OH)CH_2CH_3$

(b) $(CH_3)_2CHCH_2CH_2CH_2OH$

16.63 Write line structures for the following compounds.

(a) 3-hexanol

(b) 2,2-dimethyl-3-pentanol

16.64 Write line structures for the following compounds.

(a) 2-pentanol

(b) 3-ethyl-2-hexanol

16.65 Explain why CH_3CH_2OH has a boiling point of 78°C, but $CH_3CH_2CH_3$ has a boiling point of –42°C.

16.66 Explain why $CH_3CH_2CH_3$ is insoluble in water, but CH_3CH_2OH is soluble.

16.67 What is an ether?

16.68 How do ethers differ from alcohols?

16.69 Which of the following compounds are ethers?

(a) CH_3OCH_3

(b) CH_3CH_2OH

(c) $CH_3CH(OH)CH_3$

(d) CH_3COCH_3

(e) CH_3CH_2CHO

16.70 Which of the following compounds are ethers?

(a) CH_3CHO

(b) $(CH_3)_2C(OH)CH_3$

(c) $CH_3CH_2COCH_3$

(d) $CH_3CH_2OCH_2CH_3$

(e) $CH_3CH_2CH_2CH_2OH$

16.71 Name the following compounds.

(a) $CH_3CH_2{-}O{-}CH_2CH_2CH_3$

(b) $CH_3{-}O{-}CH_3$

16.72 Name the following compounds.

(a) $CH_3CH_2{-}O{-}CH_2CH_3$

(b) $CH_3CH_2{-}O{-}CH_3$