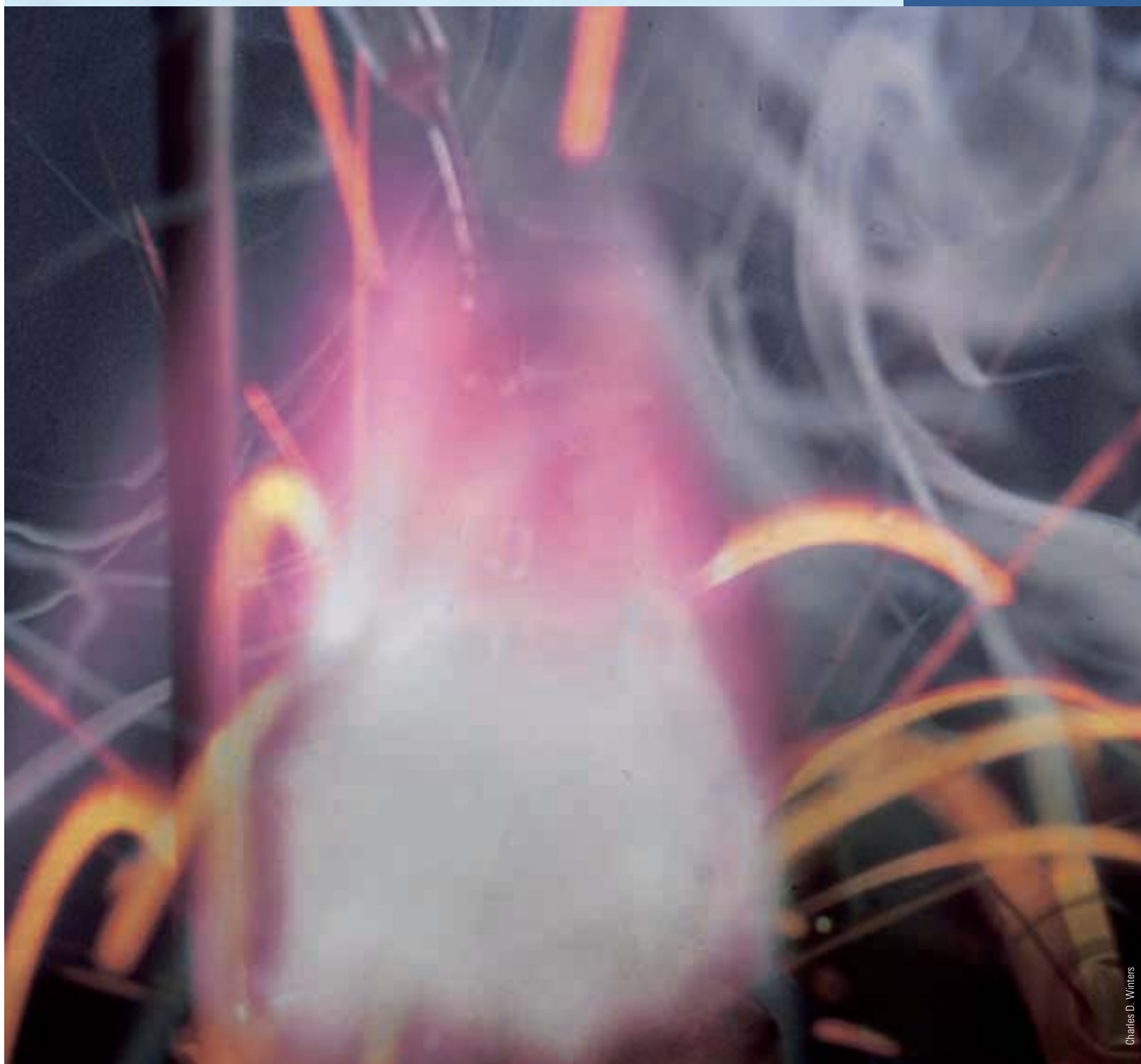


# Some Types of Chemical Reactions

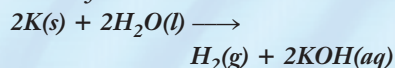
# 4





Charles D. Winters

The active metal potassium, K, reacts violently with water:



Heat released by the reaction causes unreacted potassium to give off a characteristic purple glow and ignites the hydrogen gas that is formed.

## OUTLINE

### 4-1 The Periodic Table: Metals, Nonmetals, and Metalloids

### 4-2 Aqueous Solutions: An Introduction

- 1 Electrolytes and Extent of Ionization
- 2 Strong and Weak Acids
- 3 Reversible Reactions
- 4 Strong Bases, Insoluble Bases, and Weak Bases
- 5 Solubility Guidelines for Compounds in Aqueous Solution

### 4-3 Reactions in Aqueous Solutions

### 4-4 Oxidation Numbers

### Naming Some Inorganic Compounds

### 4-5 Naming Binary Compounds

### 4-6 Naming Ternary Acids and Their Salts

### Classifying Chemical Reactions

### 4-7 Oxidation–Reduction Reactions: Introduction

### 4-8 Combination Reactions

- 1 Element + Element  $\rightarrow$  Compound
- 2 Compound + Element  $\rightarrow$  Compound

- 3 Compound + Compound  $\rightarrow$  Compound

### 4-9 Decomposition Reactions

- 1 Compound  $\rightarrow$  Element + Element
- 2 Compound  $\rightarrow$  Compound + Element
- 3 Compound  $\rightarrow$  Compound + Compound

### 4-10 Displacement Reactions

- 1  $\left[ \begin{array}{l} \text{More Active Metal +} \\ \text{Salt of Less Active} \\ \text{Metal} \end{array} \right] \longrightarrow \left[ \begin{array}{l} \text{Less Active Metal +} \\ \text{Salt of More Active Metal} \end{array} \right]$
- 2  $\left[ \begin{array}{l} \text{Active Metal +} \\ \text{Nonoxidizing Acid} \end{array} \right] \longrightarrow \left[ \begin{array}{l} \text{Hydrogen + Salt of Acid} \end{array} \right]$
- 3  $\left[ \begin{array}{l} \text{Active Nonmetal +} \\ \text{Salt of Less Active} \\ \text{Nonmetal} \end{array} \right] \longrightarrow \left[ \begin{array}{l} \text{Less Active Nonmetal +} \\ \text{Salt of More Active} \\ \text{Nonmetal} \end{array} \right]$

### 4-11 Metathesis Reactions

- 1 Acid–Base (Neutralization) Reactions: Formation of a Nonelectrolyte
- 2 Precipitation Reactions

### 4-12 Summary of Reaction Types

## OBJECTIVES

After you have studied this chapter, you should be able to

- Describe the periodic table and some of the relationships that it summarizes
- Recognize and describe nonelectrolytes, strong electrolytes, and weak electrolytes
- Recognize and classify acids (strong, weak), bases (strong, weak, insoluble), and salts (soluble, insoluble); use the solubility guidelines
- Describe reactions in aqueous solutions by writing formula unit equations, total ionic equations, and net ionic equations
- Assign oxidation numbers to elements when they are free, in compounds, or in ions
- Name and write formulas for common binary and ternary inorganic compounds
- Recognize oxidation–reduction reactions and identify which species are oxidized, reduced, oxidizing agents, and reducing agents
- Recognize and describe classes of reactions: decomposition reactions, displacement reactions, and various types of metathesis reactions

In this chapter we examine some types of chemical reactions. Millions of reactions are known, so it is useful to group them into classes, or types, so that we can deal systematically with these massive amounts of information. We describe how some compounds behave in aqueous solution, including how well their solutions conduct electricity and whether or not the compounds dissolve in water. We introduce several ways to represent chemical reactions in aqueous solution—formula unit equations, total ionic equations, and net ionic equations—and the advantages and disadvantages of these methods. We also introduce systematic ways to describe chemical reactions, as well as their reactants and their products.

Let's first take a brief look at the periodic table, which helps us to organize many properties of the elements, including their chemical reactions.

#### 4-1 THE PERIODIC TABLE: METALS, NONMETALS, AND METALLOIDS

In 1869, the Russian chemist Dmitri Mendeleev (1834–1907) and the German chemist Lothar Meyer (1830–1895) independently published arrangements of known elements that are much like the periodic table in use today. Mendeleev's classification was based primarily on chemical properties of the elements, whereas Meyer's classification was based largely on physical properties. The tabulations were surprisingly similar. Both emphasized the *periodicity*, or regular periodic repetition, of properties with increasing atomic weight.

Mendeleev arranged the known elements in order of increasing atomic weight in successive sequences so that elements with similar chemical properties fell into the same column. He noted that both physical and chemical properties of the elements vary in a periodic fashion with atomic weight. His periodic table of 1872 contained the 62 known elements (Figure 4-1). Mendeleev placed H, Li, Na, and K in his table as “Gruppe I.” These were

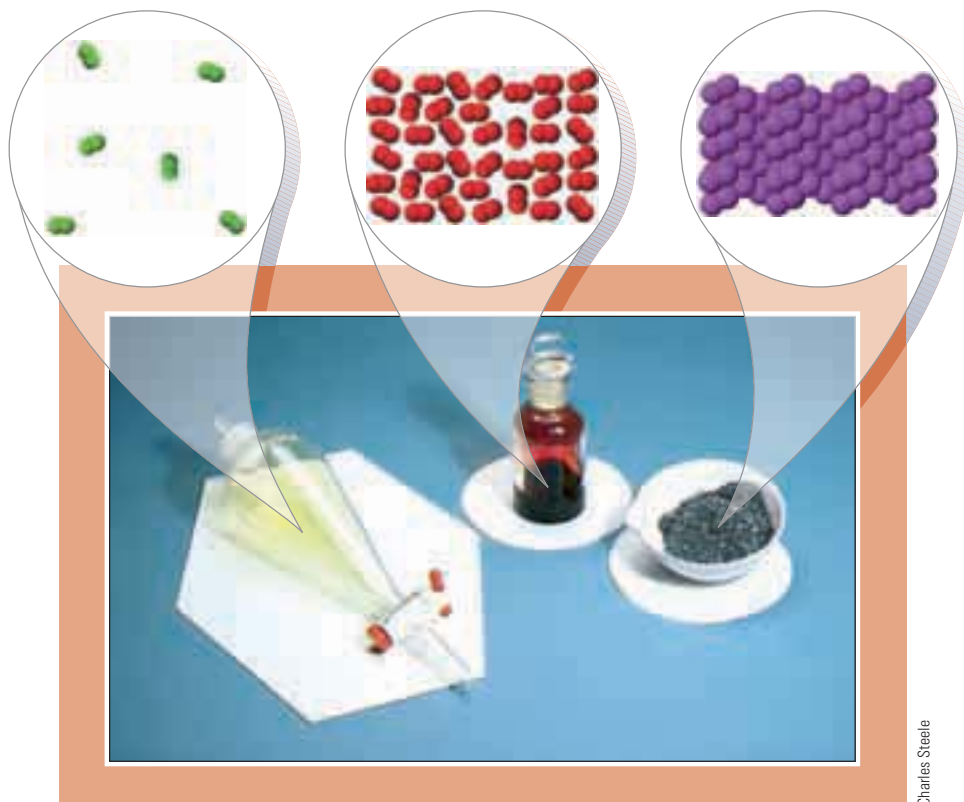
Pronounced “men-del-lay-ev.”

1A (1)		2A (2)	
3 <b>Li</b> Lithium	4 <b>Be</b> Beryllium		
11 <b>Na</b> Sodium	12 <b>Mg</b> Magnesium		
19 <b>K</b> Potassium	20 <b>Ca</b> Calcium		
37 <b>Rb</b> Rubidium	38 <b>Sr</b> Strontium		
55 <b>Cs</b> Cesium	56 <b>Ba</b> Barium		
87 <b>Fr</b> Francium	88 <b>Ra</b> Radium		

Credit source appears here

Lithium (*top left*) and sodium (*bottom left*) from Group IA. Magnesium (*top right*) and calcium (*bottom right*) from Group IIA.

known to combine with F, Cl, Br, and I of “Gruppe VII” to produce compounds that have similar formulas such as HF, LiCl, NaCl, and KI. All these compounds dissolve in water to produce solutions that conduct electricity. The “Gruppe II” elements were known to form compounds such as  $\text{BeCl}_2$ ,  $\text{MgBr}_2$ , and  $\text{CaCl}_2$ , as well as compounds with O and S from



Three of the halogens, elements from Group VIIA (left to right): chlorine, bromine, iodine.

Charles Steele

REIHEN	GRUPPE I — $\text{R}^2\text{O}$	GRUPPE II — $\text{RO}$	GRUPPE III — $\text{R}^2\text{O}^3$	GRUPPE IV $\text{RH}^4$ $\text{RO}^2$	GRUPPE V $\text{RH}^3$ $\text{R}^2\text{O}^5$	GRUPPE VI $\text{RH}^2$ $\text{RO}^3$	GRUPPE VII $\text{RH}$ $\text{R}^2\text{O}^7$	GRUPPE VIII — $\text{RO}^4$
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

**Figure 4-1** Mendeleev’s early periodic table (1872). “J” is the German symbol for iodine.



“Gruppe VI” such as MgO, CaO, MgS, and CaS. These and other chemical properties led him to devise a table in which the elements were arranged by increasing atomic weights and grouped into vertical families.

In most areas of human endeavor progress is slow and faltering. Occasionally, however, an individual develops concepts and techniques that clarify confused situations. Mendeleev was such an individual. One of the brilliant successes of his periodic table was that it provided for elements that were unknown at the time. When he encountered “missing” elements, Mendeleev left blank spaces. Some appreciation of his genius in constructing the table as he did can be gained by comparing the predicted (1871) and observed properties of germanium, which was not discovered until 1886. Mendeleev called the undiscovered element eka-silicon because it fell below silicon in his table. He was familiar with the properties of germanium’s neighboring elements. They served as the basis for his predictions of properties of germanium (Table 4-1). Some modern values for properties of germanium differ significantly from those reported in 1886. But many of the values on which Mendeleev based his predictions were also inaccurate.

Because Mendeleev’s arrangement of the elements was based on increasing *atomic weights*, several elements would have been out of place in his table. Mendeleev put the controversial elements (Te and I, Co and Ni) in locations consistent with their properties, however. He thought the apparent reversal of atomic weights was due to inaccurate values for those weights. Careful redetermination showed that the values were correct. Explanation of the locations of these “out-of-place” elements had to await the development of the concept of *atomic number*, approximately 50 years after Mendeleev’s work. The **atomic number** (Section 5-5) of an element is the number of protons in the nucleus of its atoms. (It is also the



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Not all properties of the elements can be predicted from the trends in the periodic table. For example, gallium’s melting point being low enough to melt from the heat of the hand is inconsistent with the properties of the elements above and below it.

**TABLE 4-1** Predicted and Observed Properties of Germanium

Property	Eka-Silicon Predicted, 1871	Germanium Reported, 1886	Modern Values
Atomic weight	72	72.32	72.61
Atomic volume	13 cm <sup>3</sup>	13.22 cm <sup>3</sup>	13.5 cm <sup>3</sup>
Specific gravity	5.5	5.47	5.35
Specific heat	0.073 cal/g°C	0.076 cal/g°C	0.074 cal/g°C
Maximum valence*	4	4	4
Color	Dark gray	Grayish white	Grayish white
Reaction with water	Will decompose steam with difficulty	Does not decompose water	Does not decompose water
Reactions with acids and alkalis	Slight with acids; more pronounced with alkalis	Not attacked by HCl or dilute aqueous NaOH; reacts vigorously with molten NaOH	Not dissolved by HCl or H <sub>2</sub> SO <sub>4</sub> or dilute NaOH; dissolved by concentrated NaOH
Formula of oxide	EsO <sub>2</sub>	GeO <sub>2</sub>	GeO <sub>2</sub>
Specific gravity of oxide	4.7	4.703	4.228
Specific gravity of tetrachloride	1.9 at 0°C	1.887 at 18°C	1.8443 at 30°C
Boiling point of tetrachloride	100°C	86°C	84°C
Boiling point of tetraethyl derivative	160°C	160°C	186°C

\*“Valence” refers to the combining power of a specific element.



**TABLE 4-3** *Some Physical Properties of Metals and Nonmetals*

Metals	Nonmetals
<ol style="list-style-type: none"> <li>1. High electrical conductivity that decreases with increasing temperature</li> <li>2. High thermal conductivity</li> <li>3. Metallic gray or silver luster*</li> <li>4. Almost all are solids<sup>†</sup></li> <li>5. Malleable (can be hammered into sheets)</li> <li>6. Ductile (can be drawn into wires)</li> </ol>	<ol style="list-style-type: none"> <li>1. Poor electrical conductivity (except carbon in the form of graphite)</li> <li>2. Good heat insulators (except carbon in the form of diamond)</li> <li>3. No metallic luster</li> <li>4. Solids, liquids, or gases</li> <li>5. Brittle in solid state</li> <li>6. Nonductile</li> </ol>

\*Except copper and gold.

<sup>†</sup>Except mercury; cesium and gallium melt in a protected band.

**TABLE 4-4** *Some Chemical Properties of Metals and Nonmetals*

Metals	Nonmetals
<ol style="list-style-type: none"> <li>1. Outer shells contain few electrons—usually three or fewer</li> <li>2. Form cations (positive ions) by losing electrons</li> <li>3. Form ionic compounds with nonmetals</li> <li>4. Solid state characterized by metallic bonding</li> </ol>	<ol style="list-style-type: none"> <li>1. Outer shells contain four or more electrons*</li> <li>2. Form anions (negative ions) by gaining electrons<sup>†</sup></li> <li>3. Form ionic compounds with metals<sup>†</sup> and molecular (covalent) other compounds with nonmetals</li> <li>4. Covalently bonded molecules; noble gases are monatomic</li> </ol>

\*Except hydrogen and helium.

<sup>†</sup>Except the noble gases.

elements are called **halogens**, which means “salt formers,” and the Group VIIIA elements are called **noble** (or **rare**) **gases**.

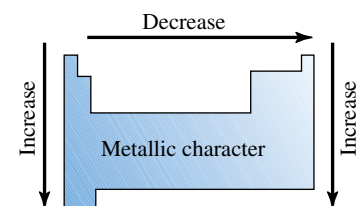
The general properties of metals and nonmetals are distinct. Physical and chemical properties that distinguish metals from nonmetals are summarized in Tables 4-3 and 4-4. Not all metals and nonmetals possess all these properties, but they share most of them to varying degrees. The physical properties of metals can be explained on the basis of metallic bonding in solids (Section 13-17).

Table 4-2, The Periodic Table, shows how we classify the known elements as *metals* (shown in blue), *nonmetals* (tan), and *metalloids* (green). The elements to the left of those touching the heavy stairstep line are metals (except hydrogen), and those to the right are nonmetals. Such a classification is somewhat arbitrary, and several elements do not fit neatly into either class. Most elements adjacent to the heavy line are often called **metalloids** (or semimetals), because they are metallic (or nonmetallic) only to a limited degree.

*Metallic character* increases from top to bottom and decreases from left to right with respect to position in the periodic table.

Cesium, atomic number 55, is the most active naturally occurring metal. Francium and radium are radioactive and do not occur in nature in appreciable amounts. Noble gases seldom bond with other elements. They are unreactive, monatomic gases. The most active nonmetal is fluorine, atomic number 9.

*Nonmetallic character* decreases from top to bottom and increases from left to right in the periodic table.

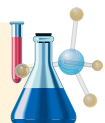


General trends in metallic character of A group elements with position in the periodic table.



Silicon, a metalloid, is widely used in the manufacture of electronic chips.

## CHEMISTRY IN USE



### The Development of Science

#### The Discovery of Phosphorus

Technology and its impact on society have always been intriguing subjects for artists. This was particularly true during the Industrial Revolution, when chemistry was on the verge of transforming itself from alchemical “magic” into a scientific discipline. It is easy to see how the scientist, toiling away in a laboratory full of strange equipment and trying to make sense of the natural world, held a certain heroic appeal to artists.

One of the more romantic accounts of chemical activity during that period is *The Alchemist in Search of the Philosophers’ Stone Discovers Phosphorus* (1771) by the English painter Joseph Wright of Derby (1734–1797). In Wright’s depiction, a stately, bearded alchemist has just isolated a new element by distillation. As the substance collects in the flask it begins to glow in the dark, illuminating the laboratory with an eerie white light and bringing the imaginary scientist to his knees in wonder. The element phosphorus was in fact named for this property—*phosphorescence*—with both words deriving from the Greek *phosphoros*, or “giving light.”

The actual discovery of elemental phosphorus was probably not quite as dramatic as Joseph Wright envisioned. It was first isolated from urine by the German chemist Henning Brand in 1669, by a much more laborious process than the one represented by the tidy distillation apparatus in Wright’s painting. The first step of the preparation, as described in a 1726 treatise entitled “Phosphoros Elemental,” in fact involved steeping 50 or 60 pails of urine in tubs for two weeks “. . . till it putrify and breed Worms”—hardly a fitting subject for eighteenth century artwork!

The glowing material was of such novelty that two of Brand’s scientific contemporaries offered to find a royal buyer for his process. Expecting a bigger reward at a later date, Brand gave the two the recipe for phosphorus in exchange for some small gifts. However, one man instead claimed the discovery for himself after repeating Brand’s work in his own laboratory. Through the other, Brand did receive a contract with the Duke of Hanover for the preparation of phosphorus; however, he was dissatisfied with his pay, and it was only after writing a number of complaint letters (and enlisting his wife to do the same) that he finally received what he felt was fair compensation for his discovery.

A number of other eighteenth-century scientific tableaux were immortalized by Wright. He was particularly fascinated



Derby Museums and Art Gallery

The Alchemist in Search of the Philosophers’ Stone Discovers Phosphorus, by Joseph Wright (1771).

by light and shadow effects. This, combined with his interest in technological subjects (the town of Derby played an important part in the beginnings of the Industrial Revolution), led him to use other unusual objects, such as glowing iron ingots (*Iron Forge*, 1772) and laboratory candles (*Experiment on a Bird in an Air Pump*, 1786), as focal points in paintings of industrial or scientific scenes.

Lisa S. Boffa  
Senior Chemist  
Exxon Corporation



Metalloids show some properties that are characteristic of both metals and nonmetals. Many of the metalloids, such as silicon, germanium, and antimony, act as semiconductors, which are important in solid-state electronic circuits. **Semiconductors** are insulators at lower temperatures but become conductors at higher temperatures (Section 13-17). The conductivities of metals, by contrast, decrease with increasing temperature.


Aluminum is the least metallic of the metals and is sometimes classified as a metalloid. It is metallic in appearance and an excellent conductor of electricity.

In this and later chapters we will study some chemical reactions of elements and their compounds and relate the reactions to the locations of the elements in the periodic table. First, we will describe some important properties of solutions and what they tell us about the nature and behavior of the dissolved substances, the solutes.

Aluminum is the most abundant metal in the earth's crust (7.5% by mass).


## 4-2 AQUEOUS SOLUTIONS: AN INTRODUCTION

Approximately three fourths of the earth's surface is covered with water. The body fluids of all plants and animals are mainly water. Thus we can see that many important chemical reactions occur in aqueous (water) solutions, or in contact with water. In Chapter 3, we introduced solutions and methods of expressing concentrations of solutions. It is useful to know the kinds of substances that are soluble in water, and the forms in which they exist, before we begin our systematic study of chemical reactions.

 **CD-ROM** Screen 5.3, Compounds in Aqueous Solution

### 1 Electrolytes and Extent of Ionization

Solutes that are water-soluble can be classified as either electrolytes or nonelectrolytes. **Electrolytes** are substances whose aqueous solutions conduct electric current. **Strong electrolytes** are substances that conduct electricity well in dilute aqueous solution. **Weak electrolytes** conduct electricity poorly in dilute aqueous solution. Aqueous solutions of **nonelectrolytes** do not conduct electricity. Electric current is carried through aqueous solution by the movement of ions. The strength of an electrolyte depends on the number of ions in solution and also on the charges on these ions (Figure 4-2).

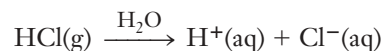
 **TOPIC:** Electrolytes/Nonelectrolytes  
**GO TO:** [www.scilinks.org](http://www.scilinks.org)  
**SCILINKS CODE:** WCH0420

**Dissociation** refers to the process in which a solid *ionic compound*, such as NaCl, separates into its ions in solution:



Recall that *ions* are charged particles. The movement of charged particles conducts electricity.

Molecular compounds, for example *pure* HCl, exist as discrete molecules and do not contain ions; however, many such compounds form ions in solution. **Ionization** refers to the process in which a *molecular compound* separates or reacts with water to form ions in solution:



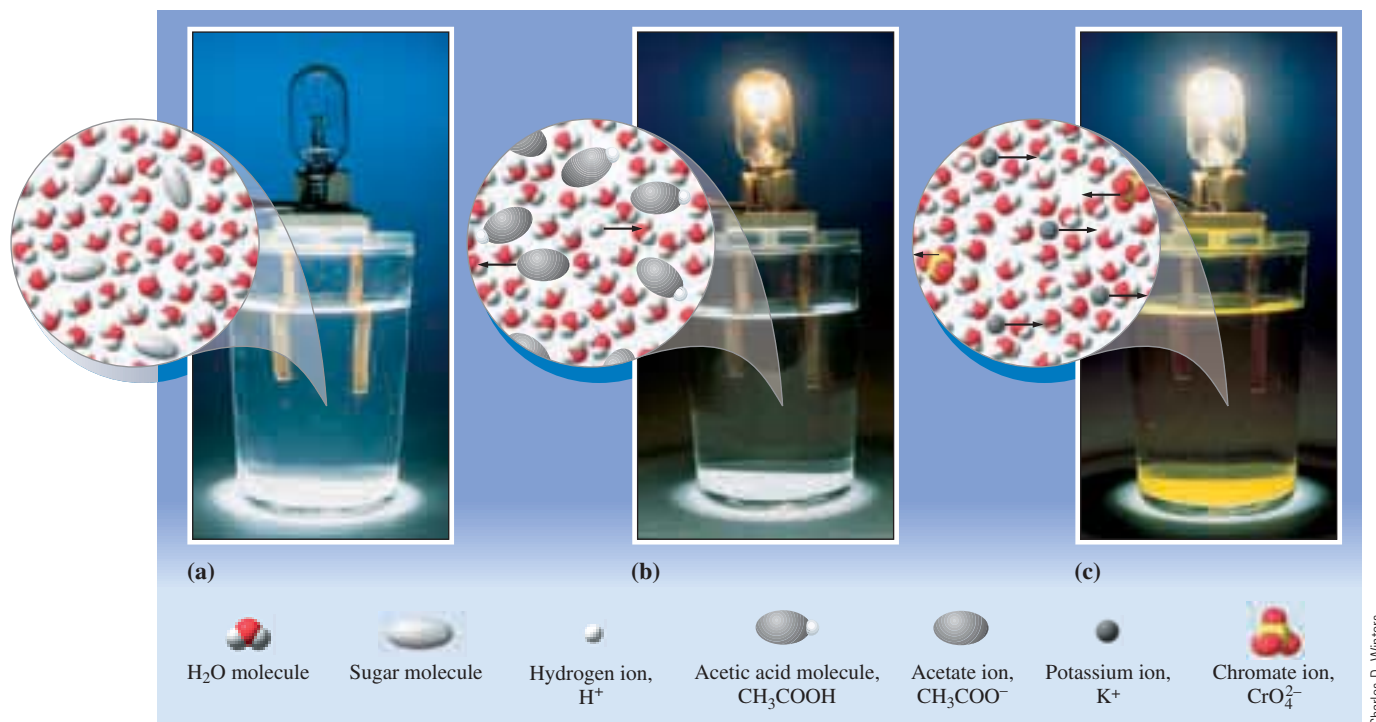
In Chapter 10, we will see that it is appropriate to represent  $\text{H}^+(\text{aq})$  as  $\text{H}_3\text{O}^+$  to emphasize its interaction with water.

Three major classes of solutes are strong electrolytes: (1) strong acids, (2) strong bases, and (3) most soluble salts. *These compounds are completely or nearly completely ionized (or dissociated) in dilute aqueous solutions*, so they are strong electrolytes.

Acids and bases are further identified in Subsections 2, 3, and 4. They are discussed in more detail in Chapter 10.

An **acid** can be defined as a substance that produces hydrogen ions,  $\text{H}^+$ , in aqueous solutions. We usually write the formulas of inorganic acids with hydrogen written first. Organic acids can often be recognized by the presence of the COOH group in the formula. A **base** is a substance that produces hydroxide ions,  $\text{OH}^-$ , in aqueous solutions. A **salt** is a compound that contains a cation other than  $\text{H}^+$  and an anion other than hydroxide ion,

Positively charged ions are called *cations*, and negatively charged ions are called *anions* (Section 2-3). The formula for a salt may include H or OH, but it *must* contain another cation *and* another anion. For example,  $\text{NaHSO}_4$  and  $\text{Al}(\text{OH})_2\text{Cl}$  are salts.



**Figure 4-2** An experiment to demonstrate the presence of ions in solution. Two copper electrodes are dipped into a liquid in a beaker. When the liquid contains significant concentrations of ions, the ions move between the electrodes to complete the circuit (which includes a light bulb). (a) Pure water and sugar are nonelectrolytes. (b) A solution of a weak electrolyte, acetic acid ( $\text{CH}_3\text{COOH}$ ); it contains low concentrations of ions, and so the bulb glows dimly. (c) A solution of a strong electrolyte, potassium chromate ( $\text{K}_2\text{CrO}_4$ ); it contains a high concentration of ions, and so the bulb glows brightly.

$\text{OH}^-$ , or oxide ion,  $\text{O}^{2-}$  (see Table 2-3). As we will see later in this chapter, salts are formed when acids react with bases.

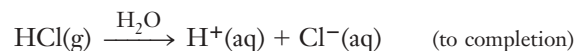
## 2 Strong and Weak Acids

Many properties of aqueous solutions of acids are due to  $\text{H}^+(\text{aq})$  ions. These are described in Section 10-4.

To give a more complete description of reactions, we indicate the physical states of reactants and products: (g) for gases, (l) for liquids, and (s) for solids. The notation (aq) following ions indicates that they are hydrated in aqueous solution; that is, they interact with water molecules in solution. The complete ionization of a strong electrolyte is indicated by a single arrow ( $\rightarrow$ ).

As a matter of convenience we place acids into two classes: strong acids and weak acids. **Strong acids** ionize (separate into hydrogen ions and stable anions) completely, or very nearly completely, in dilute aqueous solution. Seven strong acids and their anions are listed in Table 4-5. Please learn this short list; you can then assume that other acids you encounter are weak.

Because strong acids ionize completely or very nearly completely in dilute solutions, their solutions contain predominantly ions rather than acid molecules. Consider the ionization of hydrochloric acid. Pure hydrogen chloride, HCl, is a molecular compound that is a gas at room temperature and atmospheric pressure. When it dissolves in water, it reacts nearly 100% to produce a solution that contains hydrogen ions and chloride ions:



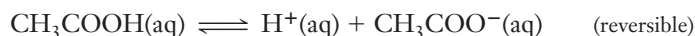
Similar equations can be written for all strong acids.

**TABLE 4-5** Some Strong Acids and Their Anions

Common Strong Acids		Anions of These Strong Acids	
Formula	Name	Formula	Name
HCl	hydrochloric acid	Cl <sup>-</sup>	chloride ion
HBr	hydrobromic acid	Br <sup>-</sup>	bromide ion
HI	hydroiodic acid	I <sup>-</sup>	iodide ion
HNO <sub>3</sub>	nitric acid	NO <sub>3</sub> <sup>-</sup>	nitrate ion
HClO <sub>4</sub>	perchloric acid	ClO <sub>4</sub> <sup>-</sup>	perchlorate ion
(HClO <sub>3</sub> )	chloric acid	ClO <sub>3</sub> <sup>-</sup>	chlorate ion
H <sub>2</sub> SO <sub>4</sub>	sulfuric acid	{ HSO <sub>4</sub> <sup>-</sup> SO <sub>4</sub> <sup>2-</sup>	hydrogen sulfate ion sulfate ion

**Weak acids** ionize only slightly (usually less than 5%) in dilute aqueous solution. Some common weak acids are listed in Appendix F. Several of them and their anions are given in Table 4-6.

The equation for the ionization of acetic acid, CH<sub>3</sub>COOH, in water is typical of weak acids:



The double arrow  $\rightleftharpoons$  generally signifies that the reaction occurs in *both* directions and that the forward reaction does not go to completion. All of us are familiar with solutions of acetic acid. Vinegar is 5% acetic acid by mass. Our use of oil and vinegar as a salad dressing suggests that acetic acid is a weak acid; we could not safely drink a 5% solution of any strong acid. To be specific, acetic acid is 0.5% ionized (and 99.5% nonionized) in 5% solution.

**TABLE 4-6** Some Common Weak Acids and Their Anions

Common Weak Acids		Anions of These Weak Acids	
Formula	Name	Formula	Name
HF*	hydrofluoric acid	F <sup>-</sup>	fluoride ion
CH <sub>3</sub> COOH	acetic acid	CH <sub>3</sub> COO <sup>-</sup>	acetate ion
HCN	hydrocyanic acid	CN <sup>-</sup>	cyanide ion
HNO <sub>2</sub> †	nitrous acid	NO <sub>2</sub> <sup>-</sup>	nitrite ion
H <sub>2</sub> CO <sub>3</sub> †	carbonic acid	{ HCO <sub>3</sub> <sup>-</sup> CO <sub>3</sub> <sup>2-</sup>	hydrogen carbonate ion carbonate ion
H <sub>2</sub> SO <sub>3</sub> †	sulfurous acid	{ HSO <sub>3</sub> <sup>-</sup> SO <sub>3</sub> <sup>2-</sup>	hydrogen sulfite ion sulfite ion
H <sub>3</sub> PO <sub>4</sub>	phosphoric acid	{ H <sub>2</sub> PO <sub>4</sub> <sup>-</sup> HPO <sub>4</sub> <sup>2-</sup> PO <sub>4</sub> <sup>3-</sup>	dihydrogen phosphate ion hydrogen phosphate ion phosphate ion
(COOH) <sub>2</sub>	oxalic acid	{ H(COO) <sub>2</sub> <sup>-</sup> (COO) <sub>2</sub> <sup>2-</sup>	hydrogen oxalate ion oxalate ion

\*HF is a weak acid, whereas HCl, HBr, and HI are strong acids.

†Free acid molecules exist only in dilute aqueous solution or not at all. Many salts of these acids are common, stable compounds, however.

Chloric acid is sometimes not listed with the common strong acids since it is not commonly encountered. However, its anion is much more common. So we have included both the acid and its anion in this table.



CD-ROM Screen 5.8, Acids

Acetic acid is the most familiar organic acid.

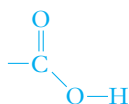
Our stomachs have linings that are much more resistant to attack by acids than are our other tissues.

The names given here correspond to the aqueous solutions.



Citrus fruits contain citric acid, so their juices are acidic. This is shown here by the red colors on the indicator paper. Acids taste sour.

The carboxylate group  $-\text{COOH}$  is



Other organic acids have other groups in the position of the  $\text{H}_3\text{C}-$  group in acetic acid.

Inorganic acids may be strong or weak.

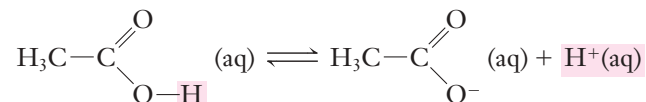


Many common food and household products are acidic (orange juice, vinegar, soft drink, citrus fruits) or basic (cleaning preparations, baking soda).

$\text{Na}^+$  and  $\text{NO}_3^-$  ions do not combine because  $\text{NaNO}_3$  is a soluble ionic compound.

Solutions of bases have a set of common properties due to the  $\text{OH}^-$  ion. These are described in Section 10-4.

A multitude of organic acids occur in living systems. Organic acids contain the carboxylate grouping of atoms,  $-\text{COOH}$ . Most common organic acids are weak. They can ionize slightly by breaking the  $\text{O}-\text{H}$  bond, as shown on the following page for acetic acid:



Organic acids are discussed in Chapter 27. Some naturally occurring organic weak acids are tartaric acid (grapes), lactic acid (sour milk), and formic acid (ants). Carbonic acid,  $\text{H}_2\text{CO}_3$ , and hydrocyanic acid,  $\text{HCN}(\text{aq})$ , are two common acids that contain carbon but that are considered to be *inorganic* acids. Inorganic acids are often called **mineral acids** because they are obtained primarily from nonliving sources.

### EXAMPLE 4-1 Strong and Weak Acids

In the following lists of common acids, which are strong and which are weak? (a)  $\text{H}_3\text{PO}_4$ ,  $\text{HCl}$ ,  $\text{H}_2\text{CO}_3$ ,  $\text{HNO}_3$ ; (b)  $\text{HClO}_4$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{HClO}$ ,  $\text{HF}$ .

#### Plan

We recall that Table 4-5 lists some common strong acids. Other *common* acids are assumed to be weak.

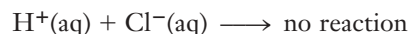
#### Solution

- (a)  $\text{HCl}$  and  $\text{HNO}_3$  are strong acids;  $\text{H}_3\text{PO}_4$  and  $\text{H}_2\text{CO}_3$  are weak acids.  
 (b)  $\text{HClO}_4$  and  $\text{H}_2\text{SO}_4$  are strong acids;  $\text{HClO}$  and  $\text{HF}$  are weak acids.

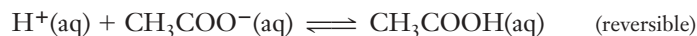
You should now work Exercises 19 and 21.

## 3 Reversible Reactions

Reactions that can occur in both directions are **reversible reactions**. We use a double arrow ( $\rightleftharpoons$ ) to indicate that a reaction is *reversible*. What is the fundamental difference between reactions that go to completion and those that are reversible? We have seen that the ionization of  $\text{HCl}$  in water is nearly complete. Suppose we dissolve some table salt,  $\text{NaCl}$ , in water and then add some dilute nitric acid to it. The resulting solution contains  $\text{Na}^+$  and  $\text{Cl}^-$  ions (from the dissociation of  $\text{NaCl}$ ) as well as  $\text{H}^+$  and  $\text{NO}_3^-$  (from the ionization of  $\text{HNO}_3$ ). The  $\text{H}^+$  and  $\text{Cl}^-$  ions do *not* react significantly to form nonionized  $\text{HCl}$  molecules; this would be the reverse of the ionization of  $\text{HCl}$ .



In contrast, when a sample of sodium acetate,  $\text{NaCH}_3\text{COO}$ , is dissolved in  $\text{H}_2\text{O}$  and mixed with nitric acid, the resulting solution initially contains  $\text{Na}^+$ ,  $\text{CH}_3\text{COO}^-$ ,  $\text{H}^+$ , and  $\text{NO}_3^-$  ions. But most of the  $\text{H}^+$  and  $\text{CH}_3\text{COO}^-$  ions combine to produce nonionized molecules of acetic acid, the reverse of the ionization of the acid. Thus, the ionization of acetic acid, like that of any other weak electrolyte, is reversible.



## 4 Strong Bases, Insoluble Bases, and Weak Bases

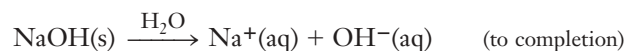
Most common bases are *ionic* metal hydroxides. **Strong bases** are soluble in water and are dissociated completely in dilute aqueous solution. The common strong bases are listed in



**TABLE 4-7** Common Strong Bases

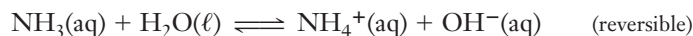
Group IA		Group IIA	
LiOH	lithium hydroxide		
NaOH	sodium hydroxide		
KOH	potassium hydroxide	Ca(OH) <sub>2</sub>	calcium hydroxide
RbOH	rubidium hydroxide	Sr(OH) <sub>2</sub>	strontium hydroxide
CsOH	cesium hydroxide	Ba(OH) <sub>2</sub>	barium hydroxide

Table 4-7. They are the hydroxides of the Group IA metals and the heavier members of Group IIA. The equation for the dissociation of sodium hydroxide in water is typical. Similar equations can be written for other strong bases.



Other metals form ionic hydroxides, but these are so sparingly soluble in water that they cannot produce strongly basic solutions. They are called **insoluble bases** or sometimes sparingly soluble bases. Typical examples include Cu(OH)<sub>2</sub>, Zn(OH)<sub>2</sub>, Fe(OH)<sub>2</sub>, and Fe(OH)<sub>3</sub>.

Common **weak bases** are molecular substances that are soluble in water but form only low concentrations of ions in solution. The most common weak base is ammonia, NH<sub>3</sub>.



Closely related N-containing compounds, the *amines*, such as methylamine, CH<sub>3</sub>NH<sub>2</sub>, and aniline, C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub>, are also weak bases. Nicotine (found in tobacco) and caffeine (found in coffee, tea, and cola drinks) are naturally occurring amines.

### EXAMPLE 4-2 Classifying Bases

From the following lists, choose (i) the strong bases, (ii) the insoluble bases, and (iii) the weak bases. (a) NaOH, Cu(OH)<sub>2</sub>, Pb(OH)<sub>2</sub>, Ba(OH)<sub>2</sub>; (b) Fe(OH)<sub>3</sub>, KOH, Mg(OH)<sub>2</sub>, Sr(OH)<sub>2</sub>, NH<sub>3</sub>.

#### Plan

(i) We recall that Table 4-7 lists the *common strong bases*. (ii) Other common metal hydroxides are assumed to be *insoluble bases*. (iii) Ammonia and closely related nitrogen-containing compounds, the amines, are the common *weak bases*.

#### Solution

- (a) (i) The strong bases are NaOH and Ba(OH)<sub>2</sub>, so  
 (ii) the insoluble bases are Cu(OH)<sub>2</sub> and Pb(OH)<sub>2</sub>.  
 (b) (i) The strong bases are KOH and Sr(OH)<sub>2</sub>, so  
 (ii) the insoluble bases are Fe(OH)<sub>3</sub> and Mg(OH)<sub>2</sub>, and  
 (iii) the weak base is NH<sub>3</sub>.

You should now work Exercises 22 and 24.



CD-ROM Screen 5.9, Bases

Strong bases are ionic compounds in the solid state.

The weak bases are *molecular substances* that dissolve in water to give slightly basic solutions; they are sometimes called *molecular bases*.



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## 5 Solubility Guidelines for Compounds in Aqueous Solution

Solubility is a complex phenomenon, and it is not possible to give a complete summary of all our observations. The following brief summary for solutes in aqueous solutions will be very useful. These generalizations are often called the *solubility guidelines*. Compounds whose solubility in water is less than about 0.02 mol/L are usually classified as insoluble

There is no sharp dividing line between “soluble” and “insoluble” compounds. Compounds whose solubilities fall near the arbitrary division are called “moderately soluble” compounds.



CD-ROM Screens 3.10, and 5.4, Solubility of Ionic Compounds



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compounds, whereas those that are more soluble are classified as soluble compounds. No gaseous or solid substances are infinitely soluble in water. You may wish to review Tables 2-3, 4-5, and 4-6. They list some common ions. Table 4-11 contains a more comprehensive list.

1. The common inorganic acids are soluble in water. Low-molecular-weight organic acids are also soluble.
2. All common compounds of the Group IA metal ions ( $\text{Li}^+$ ,  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Rb}^+$ ,  $\text{Cs}^+$ ) and the ammonium ion,  $\text{NH}_4^+$ , are soluble in water.
3. The common nitrates,  $\text{NO}_3^-$ ; acetates,  $\text{CH}_3\text{COO}^-$ ; chlorates,  $\text{ClO}_3^-$ ; and perchlorates,  $\text{ClO}_4^-$ , are soluble in water.
4. (a) The common chlorides,  $\text{Cl}^-$ , are soluble in water except  $\text{AgCl}$ ,  $\text{Hg}_2\text{Cl}_2$ , and  $\text{PbCl}_2$ .  
(b) The common bromides,  $\text{Br}^-$ , and iodides,  $\text{I}^-$ , show approximately the same solubility behavior as chlorides, but there are some exceptions. As these halide ions ( $\text{Cl}^-$ ,  $\text{Br}^-$ ,  $\text{I}^-$ ) increase in size, the solubilities of their slightly soluble compounds decrease.  
(c) The common fluorides,  $\text{F}^-$ , are soluble in water except  $\text{MgF}_2$ ,  $\text{CaF}_2$ ,  $\text{SrF}_2$ ,  $\text{BaF}_2$ , and  $\text{PbF}_2$ .
5. The common sulfates,  $\text{SO}_4^{2-}$ , are soluble in water except  $\text{PbSO}_4$ ,  $\text{BaSO}_4$ , and  $\text{HgSO}_4$ ;  $\text{CaSO}_4$ ,  $\text{SrSO}_4$ , and  $\text{Ag}_2\text{SO}_4$  are moderately soluble.
6. The common metal hydroxides,  $\text{OH}^-$ , are *insoluble* in water except those of the Group IA metals and the heavier members of the Group IIA metals, beginning with  $\text{Ca}(\text{OH})_2$ .
7. The common carbonates,  $\text{CO}_3^{2-}$ , phosphates,  $\text{PO}_4^{3-}$ , and arsenates,  $\text{AsO}_4^{3-}$ , are *insoluble* in water except those of the Group IA metals and  $\text{NH}_4^+$ .  $\text{MgCO}_3$  is moderately soluble.
8. The common sulfides,  $\text{S}^{2-}$ , are *insoluble* in water except those of the Group IA and Group IIA metals and the ammonium ion.

Table 4-8 summarizes much of the information about the solubility guidelines.

We have distinguished between strong and weak electrolytes and between soluble and insoluble compounds. Let us now see how we can describe chemical reactions in aqueous solutions.

### EXAMPLE 4-3 Solubility of Some Common Ionic Salts

From the following compounds, choose (a) those that are likely to be soluble in water and (b) those that are likely to be insoluble:  $\text{NaBr}$ ,  $\text{Cu}(\text{OH})_2$ ,  $\text{PbCl}_2$ ,  $\text{AgI}$ ,  $\text{Fe}_2\text{O}_3$ ,  $\text{Mg}(\text{NO}_3)_2$ ,  $(\text{NH}_4)_2\text{SO}_4$ .

#### Plan

We recall from the guidelines and Table 4-8 that all  $\text{Na}^+$ ,  $\text{K}^+$ , and  $\text{NH}_4^+$  salts are soluble. Therefore, we predict that  $\text{NaBr}$  and  $(\text{NH}_4)_2\text{SO}_4$  will be soluble. Similarly,  $\text{NO}_3^-$  salts are soluble, so  $\text{Mg}(\text{NO}_3)_2$  should also be soluble. All other compounds listed in this example should be insoluble.

#### Solution

- (a) The soluble compounds are  $\text{NaBr}$ ,  $(\text{NH}_4)_2\text{SO}_4$ , and  $\text{Mg}(\text{NO}_3)_2$ .  
(b) The insoluble compounds are  $\text{Cu}(\text{OH})_2$ ,  $\text{PbCl}_2$ ,  $\text{AgI}$ , and  $\text{Fe}_2\text{O}_3$ .

You should now work Exercises 27 and 30.

**TABLE 4-8** Solubility Guidelines for Common Ionic Compounds in Water

Generally Soluble	Exceptions
Na <sup>+</sup> , K <sup>+</sup> , NH <sub>4</sub> <sup>+</sup> compounds	No common exceptions
fluorides (F <sup>-</sup> )	Insoluble: MgF <sub>2</sub> , CaF <sub>2</sub> , SrF <sub>2</sub> , BaF <sub>2</sub> , PbF <sub>2</sub>
chlorides (Cl <sup>-</sup> )	Insoluble: AgCl, Hg <sub>2</sub> Cl <sub>2</sub> Soluble in hot water: PbCl <sub>2</sub>
bromides (Br <sup>-</sup> )	Insoluble: AgBr, Hg <sub>2</sub> Br <sub>2</sub> , PbBr <sub>2</sub> Moderately soluble: HgBr <sub>2</sub>
iodides (I <sup>-</sup> )	Insoluble: many heavy-metal iodides
sulfates (SO <sub>4</sub> <sup>2-</sup> )	Insoluble: BaSO <sub>4</sub> , PbSO <sub>4</sub> , HgSO <sub>4</sub> Moderately soluble: CaSO <sub>4</sub> , SrSO <sub>4</sub> , Ag <sub>2</sub> SO <sub>4</sub>
nitrates (NO <sub>3</sub> <sup>-</sup> ), nitrites (NO <sub>2</sub> <sup>-</sup> )	Moderately soluble: AgNO <sub>2</sub>
chlorates (ClO <sub>3</sub> <sup>-</sup> ), perchlorates (ClO <sub>4</sub> <sup>-</sup> )	No common exceptions
acetates (CH <sub>3</sub> COO <sup>-</sup> )	Moderately soluble: AgCH <sub>3</sub> COO
Generally Insoluble	Exceptions
sulfides (S <sup>2-</sup> ) <sup>†</sup>	Soluble: those of NH <sub>4</sub> <sup>+</sup> , Na <sup>+</sup> , K <sup>+</sup> , Mg <sup>2+</sup> , Ca <sup>2+</sup>
oxides (O <sup>2-</sup> ), hydroxides (OH <sup>-</sup> )	Soluble: Li <sub>2</sub> O*, LiOH, Na <sub>2</sub> O*, NaOH, K <sub>2</sub> O*, KOH, BaO*, Ba(OH) <sub>2</sub> Moderately soluble: CaO*, Ca(OH) <sub>2</sub> , SrO*, Sr(OH) <sub>2</sub>
carbonates (CO <sub>3</sub> <sup>2-</sup> ), phosphates (PO <sub>4</sub> <sup>3-</sup> ), arsenates (AsO <sub>4</sub> <sup>3-</sup> )	Soluble: those of NH <sub>4</sub> <sup>+</sup> , Na <sup>+</sup> , K <sup>+</sup>

\* Dissolves with evolution of heat and formation of hydroxides.

† Dissolves with formation of HS<sup>-</sup> and H<sub>2</sub>S.

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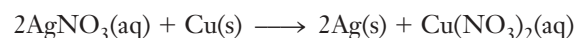
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### 4-3 REACTIONS IN AQUEOUS SOLUTIONS

Many important chemical reactions occur in aqueous solutions. In this chapter you should learn to describe such aqueous reactions and to predict the products of many reactions.

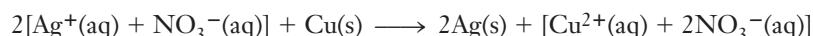
First, let's look at how we write chemical equations that describe reactions in aqueous solutions. We use three kinds of chemical equations. Table 4-9 shows the kinds of information about each substance that we use in writing equations for reactions in aqueous solutions. Some typical examples are included. Refer to Table 4-9 often as you study the following sections.

1. In **formula unit equations**, we show complete formulas for all compounds. When metallic copper is added to a solution of (colorless) silver nitrate, the more active metal — copper — displaces silver ions from the solution. The resulting solution contains blue copper(II) nitrate, and metallic silver forms as a finely divided solid (Figure 4-3):



Both silver nitrate and copper(II) nitrate are soluble ionic compounds (for solubility guidelines see Table 4-8 and Section 4-2 Part 5).

2. In **total ionic equations**, formulas are written to show the (predominant) form in which each substance exists when it is in contact with aqueous solution. We often use brackets in total ionic equations to show ions that have a common source or that remain in solution after the reaction is complete. The total ionic equation for this reaction is



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The White Cliffs of Dover, England, are composed mainly of calcium carbonate (CaCO<sub>3</sub>).

Because we have not studied periodic trends in properties of transition metals, it would be difficult for you to predict that Cu is more active than Ag. The fact that this reaction occurs (see Figure 4-3) shows that it is.

**TABLE 4-9** Bonding, Solubility, Electrolyte Characteristics, and Predominant Forms of Solutes in Contact with Water

	Acids		Bases			Salts	
	Strong acids	Weak acids	Strong bases	Insoluble bases	Weak bases	Soluble salts	Insoluble salts
Examples	HCl HNO <sub>3</sub>	CH <sub>3</sub> COOH HF	NaOH Ca(OH) <sub>2</sub>	Mg(OH) <sub>2</sub> Al(OH) <sub>3</sub>	NH <sub>3</sub> CH <sub>3</sub> NH <sub>2</sub>	KCl, NaNO <sub>3</sub> , NH <sub>4</sub> Br	BaSO <sub>4</sub> , AgCl, Ca <sub>3</sub> (PO <sub>4</sub> ) <sub>2</sub>
Pure compound ionic or molecular?	Molecular	Molecular	Ionic	Ionic	Molecular	Ionic	Ionic
Water-soluble or insoluble?	Soluble*	Soluble*	Soluble	Insoluble	Soluble <sup>†</sup>	Soluble	Insoluble
≈ 100% ionized or dissociated in dilute aqueous solution?	Yes	No	Yes	(footnote <sup>‡</sup> )	No	Yes <sup>§</sup>	(footnote <sup>‡</sup> )
Written in ionic equations as	Separate ions	Molecules	Separate ions	Complete formulas	Molecules	Separate ions	Complete formulas

\*Most common inorganic acids and the low-molecular-weight organic acids (—COOH) are water-soluble.

<sup>†</sup>The low-molecular-weight amines are water-soluble.

<sup>‡</sup>The very small concentrations of “insoluble” metal hydroxides and insoluble salts in saturated aqueous solutions are nearly completely dissociated.

<sup>§</sup>There are a few exceptions. A few soluble salts are molecular (and not ionic) compounds.

Ions that appear in solution on both sides of the total ionic equation are called **spectator ions**; they undergo no change in the chemical reaction.

Brackets are not used in net ionic equations.

This is why it is important to know how and when to construct net ionic equations from formula unit equations.

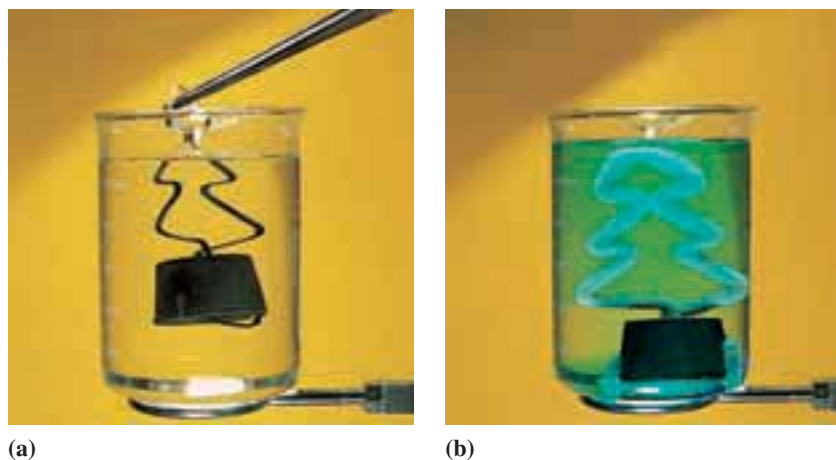
Examination of the total ionic equation shows that NO<sub>3</sub><sup>−</sup> ions do not participate in the reaction. Because they do not change, they are often called “spectator” ions.

3. In **net ionic equations**, we show only the species that react. The net ionic equation is obtained by eliminating the spectator ions and the brackets from the total ionic equation.



Net ionic equations allow us to focus on the *essence* of a chemical reaction in aqueous solutions. On the other hand, if we are dealing with stoichiometric calculations, we frequently must deal with formula weights and therefore with the *complete* formulas of all species. In

**Figure 4-3** (a) Copper wire and a silver nitrate solution. (b) The copper wire has been placed in the solution and some finely divided silver has deposited on the wire. The solution is blue because it contains copper(II) nitrate.



Charles Steele

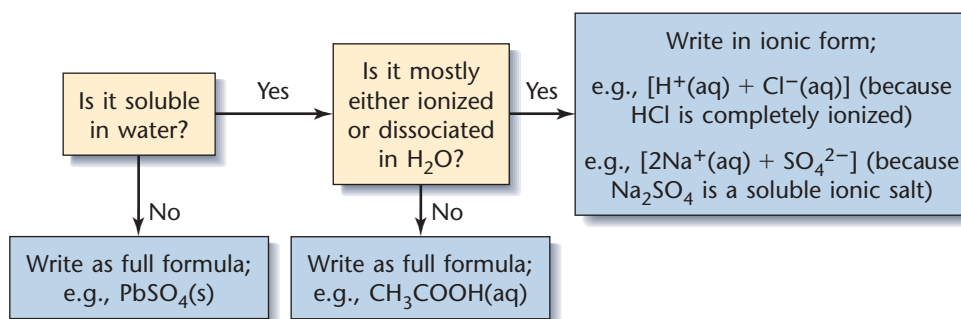


 **Problem-Solving Tip: Writing Ionic Equations**

The following chart will help in deciding which formula units are to be written as separate ions in the total ionic equation and which ones are to be written as unchanged formula units. You must answer two questions about a substance to determine whether it should be written in ionic form or as a formula unit in the total and net ionic equations.

1. Does it dissolve in water? If not, write the full formula.
2. (a) If it dissolves, does it ionize (a strong acid)?  
(b) If it dissolves, does it dissociate (a strong base or a soluble salt)?

If the answer to *either part* of the second question is yes, the substance is a soluble strong electrolyte, and its formula is written in ionic form.



The only common substances that should be written as ions in ionic equations are (1) strong acids, (2) strong bases, and (3) soluble ionic salts.



CD-ROM Screen 5.5, Equations of Reaction in Aqueous Solution: Net Ionic Equations

such cases, formula unit equations are more useful. Total ionic equations provide the bridge between the two.

Recall the lists of strong acids (Table 4-5) and strong bases (Table 4-7). These acids and bases are completely or almost completely ionized or dissociated in dilute aqueous solutions. Other common acids and bases are either insoluble or only slightly ionized or dissociated. In addition, the solubility guidelines (Table 4-8 and Section 4-2 Part 5) allow you to determine which salts are soluble in water. Most salts that are soluble in water are also strong electrolytes. Exceptions such as lead acetate,  $\text{Pb}(\text{CH}_3\text{COO})_2$ , which is soluble but does not ionize appreciably, will be noted as they are encountered.

## 4-4 OXIDATION NUMBERS

The many reactions that involve the transfer of electrons from one species to another are called **oxidation–reduction reactions**, or simply, **redox reactions**. We use oxidation numbers to keep track of electron transfers. The systematic naming of compounds (Sections 4-5 and 4-6) also makes use of oxidation numbers.

The **oxidation number**, or **oxidation state**, of an element in a simple *binary* ionic compound is the number of electrons gained or lost by an atom of that element when it forms the compound. In the case of a single-atom ion, it corresponds to the actual charge on the ion. In molecular compounds, oxidation numbers do not have the same significance they

Oxidation–reduction and displacement reactions are discussed in Sections 4-7 and 4-10.

Binary means two. Binary compounds contain two elements.

have in binary ionic compounds. Oxidation numbers, however, are very useful aids in writing formulas and in balancing equations. In molecular species, the oxidation numbers are assigned according to an arbitrary set of rules. The element farther to the right and higher up in the periodic table is assigned a negative oxidation number, and the element farther to the left and lower down in the periodic table is assigned a positive oxidation number.

Some rules for assigning oxidation numbers follow. These rules are not comprehensive, but they cover most cases. In applying these rules, keep in mind two important points. First, oxidation numbers are always assigned on a per atom basis; second, treat the rules in order of *decreasing* importance—the first rule that applies takes precedence over any subsequent rules that seem to apply.

Polyatomic elements have two or more atoms per molecule.



CD-ROM Screen 5.13, Oxidation Numbers

1. The oxidation number of the atoms in any free, uncombined element is zero. This includes polyatomic elements such as  $\text{H}_2$ ,  $\text{O}_2$ ,  $\text{O}_3$ ,  $\text{P}_4$ , and  $\text{S}_8$ .
2. The oxidation number of an element in a simple (monatomic) ion is equal to the charge on the ion.
3. The *sum* of the oxidation numbers of all atoms in a compound is zero.
4. In a polyatomic ion, the *sum* of the oxidation numbers of the constituent atoms is equal to the charge on the ion.
5. Fluorine has an oxidation number of  $-1$  in its compounds.
6. Hydrogen has an oxidation number of  $+1$  in compounds unless it is combined with metals, in which case it has an oxidation number of  $-1$ . Examples of these exceptions are  $\text{NaH}$  and  $\text{CaH}_2$ .
7. Oxygen usually has an oxidation number of  $-2$  in its compounds. There are some exceptions:
  - a. Oxygen has an oxidation number of  $-1$  in hydrogen peroxide,  $\text{H}_2\text{O}_2$ , and in peroxides, which contain the  $\text{O}_2^{2-}$  ion; examples are  $\text{CaO}_2$  and  $\text{Na}_2\text{O}_2$ .
  - b. Oxygen has an oxidation number of  $-\frac{1}{2}$  in superoxides, which contain the  $\text{O}_2^-$  ion; examples are  $\text{KO}_2$  and  $\text{RbO}_2$ .
  - c. When combined with fluorine in  $\text{OF}_2$ , oxygen has an oxidation number of  $+2$ .
8. The position of the element in the periodic table helps to assign its oxidation number:
  - a. Group IA elements have oxidation numbers of  $+1$  in all of their compounds.
  - b. Group IIA elements have oxidation numbers of  $+2$  in all of their compounds.
  - c. Group IIIA elements have oxidation numbers of  $+3$  in all of their compounds, with a few rare exceptions.
  - d. Group VA elements have oxidation numbers of  $-3$  in *binary* compounds with metals, with H, or with  $\text{NH}_4^+$ . Exceptions are compounds with a Group VA element combined with an element to its right in the periodic table; in this case, their oxidation numbers can be found by using rules 3 and 4.
  - e. Group VIA elements below oxygen have oxidation numbers of  $-2$  in *binary* compounds with metals, with H, or with  $\text{NH}_4^+$ . When these elements are combined with oxygen or with a lighter halogen, their oxidation numbers can be found by using rules 3 and 4.

**TABLE 4-10** Common Oxidation Numbers (States) for Group A Elements in Compounds and Ions

Element(s)	Common Ox. Nos.	Examples	Other Ox. Nos.
H	+1	H <sub>2</sub> O, CH <sub>4</sub> , NH <sub>4</sub> Cl	-1 in metal hydrides, e.g., NaH, CaH <sub>2</sub>
Group IA	+1	KCl, NaH, RbNO <sub>3</sub> , K <sub>2</sub> SO <sub>4</sub>	None
Group IIA	+2	CaCl <sub>2</sub> , MgH <sub>2</sub> , Ba(NO <sub>3</sub> ) <sub>2</sub> , SrSO <sub>4</sub>	None
Group IIIA	+3	AlCl <sub>3</sub> , BF <sub>3</sub> , Al(NO <sub>3</sub> ) <sub>3</sub> , GaI <sub>3</sub>	None in common compounds
Group IVA	+2 +4	CO, PbO, SnCl <sub>2</sub> , Pb(NO <sub>3</sub> ) <sub>2</sub> CCl <sub>4</sub> , SiO <sub>2</sub> , SiO <sub>3</sub> <sup>2-</sup> , SnCl <sub>4</sub>	Many others are also seen for C and Si
Group VA	-3 in binary compounds with metals -3 in NH <sub>4</sub> <sup>+</sup> , binary compounds with H	Mg <sub>3</sub> N <sub>2</sub> , Na <sub>3</sub> P, Cs <sub>3</sub> As NH <sub>3</sub> , PH <sub>3</sub> , AsH <sub>3</sub> , NH <sub>4</sub> <sup>+</sup>	+3, e.g., NO <sub>2</sub> <sup>-</sup> , PCl <sub>3</sub> +5, e.g., NO <sub>3</sub> <sup>-</sup> , PO <sub>4</sub> <sup>3-</sup> , AsF <sub>5</sub> , P <sub>4</sub> O <sub>10</sub>
O	-2	H <sub>2</sub> O, P <sub>4</sub> O <sub>10</sub> , Fe <sub>2</sub> O <sub>3</sub> , CaO, ClO <sub>3</sub> <sup>-</sup>	+2 in OF <sub>2</sub> -1 in peroxides, e.g., H <sub>2</sub> O <sub>2</sub> , Na <sub>2</sub> O <sub>2</sub> -½ in superoxides, e.g., KO <sub>2</sub> , RbO <sub>2</sub>
Group VIA (other than O)	-2 in binary compounds with metals and H -2 in binary compounds with NH <sub>4</sub> <sup>+</sup>	H <sub>2</sub> S, CaS, Fe <sub>2</sub> S <sub>3</sub> , Na <sub>2</sub> Se (NH <sub>4</sub> ) <sub>2</sub> S, (NH <sub>4</sub> ) <sub>2</sub> Se	+4 with O and the lighter halogens, e.g., SO <sub>2</sub> , SeO <sub>2</sub> , Na <sub>2</sub> SO <sub>3</sub> , SO <sub>3</sub> <sup>2-</sup> , SF <sub>4</sub> +6 with O and the lighter halogens, e.g., SO <sub>3</sub> , TeO <sub>3</sub> , H <sub>2</sub> SO <sub>4</sub> , SO <sub>4</sub> <sup>2-</sup> , SF <sub>6</sub>
Group VIIA	-1 in binary compounds with metals and H -1 in binary compounds with NH <sub>4</sub> <sup>+</sup>	MgF <sub>2</sub> , KI, ZnCl <sub>2</sub> , FeBr <sub>3</sub> NH <sub>4</sub> Cl, NH <sub>4</sub> Br	Cl, Br, or I with O or with a lighter halogen +1, e.g., BrF, ClO <sup>-</sup> , BrO <sup>-</sup> +3, e.g., ICl <sub>3</sub> , ClO <sub>2</sub> <sup>-</sup> , BrO <sub>2</sub> <sup>-</sup> +5, e.g., BrF <sub>5</sub> , ClO <sub>3</sub> <sup>-</sup> , BrO <sub>3</sub> <sup>-</sup> +7, e.g., IF <sub>7</sub> , ClO <sub>4</sub> <sup>-</sup> , BrO <sub>4</sub> <sup>-</sup>

- f. Group VIIA elements have oxidation numbers of  $-1$  in *binary* compounds with metals, with H, with NH<sub>4</sub><sup>+</sup>, or with a heavier halogen. When these elements except fluorine (i.e., Cl, Br, I) are combined with oxygen or with a lighter halogen, their oxidation numbers can be found by using rules 3 and 4.

Table 4-10 summarizes rules 5 through 8, with many examples.

#### EXAMPLE 4-4 Oxidation Numbers

Determine the oxidation numbers of nitrogen in the following species: (a) N<sub>2</sub>O<sub>4</sub>, (b) NH<sub>3</sub>, (c) HNO<sub>3</sub>, (d) NO<sub>3</sub><sup>-</sup>, (e) N<sub>2</sub>.

##### Plan

We first assign oxidation numbers to elements that exhibit a single common oxidation number (see Table 4-10). We recall that oxidation numbers are represented *per atom*, that the sum of the oxidation numbers in a compound is zero, and that the sum of the oxidation numbers in an ion equals the charge on the ion.



Charles Steele

Aqueous solutions of some compounds that contain chromium. *Left to right:* chromium(II) chloride, CrCl<sub>2</sub>, is blue; chromium(III) chloride, CrCl<sub>3</sub>, is green; potassium chromate, K<sub>2</sub>CrO<sub>4</sub>, is yellowish; potassium dichromate, K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>, is orangish.

By convention, *oxidation numbers* are represented as  $+n$  and  $-n$ , but ionic charges are represented as  $n+$  and  $n-$ . We shall circle oxidation numbers associated with formulas and show them in red. Both oxidation numbers and ionic charges can be combined algebraically.

Usually the element with the positive oxidation number is written first. For historic reasons, however, in compounds containing nitrogen and hydrogen, such as  $\text{NH}_3$ , and many compounds containing carbon and hydrogen, such as  $\text{CH}_4$ , hydrogen is written last, although it has a positive oxidation number.

**Solution**

- (a) The oxidation number of O is  $-2$ . The sum of the oxidation numbers for all atoms in a compound must be zero:

$$\begin{array}{l} \text{ox. no./atom: } \begin{array}{c} \textcircled{x} \quad \textcircled{-2} \\ \text{N}_2\text{O}_4 \end{array} \\ \text{total ox. no.: } 2x + 4(-2) = 0 \quad \text{or} \quad x = \textcircled{+4} \end{array}$$

- (b) The oxidation number of H is  $+1$ :

$$\begin{array}{l} \text{ox. no./atom: } \begin{array}{c} \textcircled{x} \quad \textcircled{+1} \\ \text{NH}_3 \end{array} \\ \text{total ox. no.: } x + 3(1) = 0 \quad \text{or} \quad x = \textcircled{-3} \end{array}$$

- (c) The oxidation number of H is  $+1$  and the oxidation number of O is  $-2$ .

$$\begin{array}{l} \text{ox. no./atom: } \begin{array}{c} \textcircled{+1} \quad \textcircled{x} \quad \textcircled{-2} \\ \text{HNO}_3 \end{array} \\ \text{total ox. no.: } 1 + x + 3(-2) = 0 \quad \text{or} \quad x = \textcircled{+5} \end{array}$$

- (d) The sum of the oxidation numbers for all atoms in an ion equals the charge on the ion:

$$\begin{array}{l} \text{ox. no./atom: } \begin{array}{c} \textcircled{x} \quad \textcircled{-2} \\ \text{NO}_3^- \end{array} \\ \text{total ox. no.: } x + 3(-2) = -1 \quad \text{or} \quad x = \textcircled{+5} \end{array}$$

- (e) The oxidation number of any free element is **zero**.

*You should now work Exercise 42.*

## NAMING SOME INORGANIC COMPOUNDS



CD-ROM Screen 3.11, Naming Ionic Compounds

The rules for naming inorganic compounds are set down by the Committee on Inorganic Nomenclature of the International Union of Pure and Applied Chemistry (IUPAC). The names and formulas of a few organic compounds were given in Table 2-2, and more systematic rules for naming them will appear in Chapter 27.

### 4-5 NAMING BINARY COMPOUNDS

Binary compounds consist of two elements; they may be either ionic or molecular. The rule is to name the more metallic element first and the less metallic element second. The less metallic element is named by adding an “-ide” suffix to the element’s *unambiguous* stem. Stems for the nonmetals follow.

IIIA		IVA		VA		VIA		VIIA	
B	bor	C	carb	N	nitr	O	ox	H	hydr
		Si	silic	P	phosph	S	sulf	F	fluor
				As	arsen	Se	selen	Cl	chlor
				Sb	antimon	Te	tellur	Br	brom
								I	iod

**Binary ionic compounds** contain metal cations and nonmetal anions. The cation is named first and the anion second.

Millions of compounds are known, so it is important to be able to associate names and formulas in a systematic way.

The stem for each element is derived from the name of the element.



Formula	Name	Formula	Name
KBr	potassium bromide	Rb <sub>2</sub> S	rubidium sulfide
CaCl <sub>2</sub>	calcium chloride	Ba <sub>3</sub> N <sub>2</sub>	barium nitride
NaH	sodium hydride	SrO	strontium oxide

Notice that there is a space between the name of the cation and the name of the anion.

The preceding method is sufficient for naming binary ionic compounds containing metals that exhibit *only one oxidation number* other than zero (Section 4-4). Most transition metals and the metals of Groups IIIA (except Al), IVA, and VA exhibit more than one oxidation number. These metals may form two or more binary compounds with the same nonmetal. To distinguish among all the possibilities, the oxidation number of the metal is indicated by a Roman numeral in parentheses following its name. This method can be applied to any binary compound of a metal and a nonmetal.

Roman numerals are *not* necessary for metals that commonly exhibit only one oxidation number in their compounds.

Formula	Ox. No. of Metal	Name	Formula	Ox. No. of Metal	Name
Cu <sub>2</sub> O	+1	copper(I) oxide	SnCl <sub>2</sub>	+2	tin(II) chloride
CuF <sub>2</sub>	+2	copper(II) fluoride	SnCl <sub>4</sub>	+4	tin(IV) chloride
FeS	+2	iron(II) sulfide	PbO	+2	lead(II) oxide
Fe <sub>2</sub> O <sub>3</sub>	+3	iron(III) oxide	PbO <sub>2</sub>	+4	lead(IV) oxide

Notice that prefixes are not used to indicate the number of ions in the formula. For example, the correct name is tin(IV) chloride not tetratin monochloride.

The advantage of the IUPAC system is that if you know the formula, you can write the exact and unambiguous name; if you are given the name, you can write the formula at once. An older method, still in use but not recommended by the IUPAC, uses “-ous” and “-ic” suffixes to indicate lower and higher oxidation numbers, respectively. This system can distinguish between only two different oxidation numbers for a metal. It is therefore not as useful as the Roman numeral system.

Familiarity with the older system is still necessary. It is still widely used in many scientific, engineering, and medical fields.

Formula	Ox. No. of Metal	Name	Formula	Ox. No. of Metal	Name
CuCl	+1	cuprous chloride	SnF <sub>2</sub>	+2	stannous fluoride
CuCl <sub>2</sub>	+2	cupric chloride	SnF <sub>4</sub>	+4	stannic fluoride
FeO	+2	ferrous oxide	Hg <sub>2</sub> Cl <sub>2</sub>	+1	mercurous chloride
FeBr <sub>3</sub>	+3	ferric bromide	HgCl <sub>2</sub>	+2	mercuric chloride

Some compounds contain polyatomic ions that behave much like monatomic anions. Compounds that contain these ions are called **pseudobinary ionic compounds**. The prefix “pseudo-” means “false”; these compounds are named as though they were binary compounds. The common examples of such polyatomic anions are the hydroxide ion, OH<sup>-</sup>, and the cyanide ion, CN<sup>-</sup>. The ammonium ion, NH<sub>4</sub><sup>+</sup>, is the common cation that behaves like a simple metal cation.

Formula	Name	Formula	Name
NH <sub>4</sub> I	ammonium iodide	NH <sub>4</sub> CN	ammonium cyanide
Ca(CN) <sub>2</sub>	calcium cyanide	Cu(OH) <sub>2</sub>	copper(II) hydroxide or cupric hydroxide
NaOH	sodium hydroxide	Fe(OH) <sub>3</sub>	iron(III) hydroxide or ferric hydroxide

A list of common cations and anions appears in Table 4-11. It will enable you to name many of the ionic compounds you encounter.

Nearly all **binary molecular compounds** involve two *nonmetals* bonded together.

**TABLE 4-11** Formulas, Ionic Charges, and Names for Some Common Ions

Common Cations			Common Anions		
Formula	Charge	Name	Formula	Charge	Name
Li <sup>+</sup>	1+	lithium ion	F <sup>-</sup>	1-	fluoride ion
Na <sup>+</sup>	1+	sodium ion	Cl <sup>-</sup>	1-	chloride ion
K <sup>+</sup>	1+	potassium ion	Br <sup>-</sup>	1-	bromide ion
NH <sub>4</sub> <sup>+</sup>	1+	ammonium ion	I <sup>-</sup>	1-	iodide ion
Ag <sup>+</sup>	1+	silver ion	OH <sup>-</sup>	1-	hydroxide ion
Mg <sup>2+</sup>	2+	magnesium ion	CN <sup>-</sup>	1-	cyanide ion
Ca <sup>2+</sup>	2+	calcium ion	ClO <sup>-</sup>	1-	hypochlorite ion
Ba <sup>2+</sup>	2+	barium ion	ClO <sub>2</sub> <sup>-</sup>	1-	chlorite ion
Cd <sup>2+</sup>	2+	cadmium ion	ClO <sub>3</sub> <sup>-</sup>	1-	chlorate ion
Zn <sup>2+</sup>	2+	zinc ion	ClO <sub>4</sub> <sup>-</sup>	1-	perchlorate ion
Cu <sup>2+</sup>	2+	copper(II) ion or cupric ion	CH <sub>3</sub> COO <sup>-</sup>	1-	acetate ion
Hg <sub>2</sub> <sup>2+</sup>	2+	mercury(I) ion or mercurous ion	MnO <sub>4</sub> <sup>-</sup>	1-	permanganate ion
Hg <sup>2+</sup>	2+	mercury(II) ion or mercuric ion	NO <sub>2</sub> <sup>-</sup>	1-	nitrite ion
Mn <sup>2+</sup>	2+	manganese(II) ion or manganous ion	NO <sub>3</sub> <sup>-</sup>	1-	nitrate ion
Co <sup>2+</sup>	2+	cobalt(II) ion or cobaltous ion	SCN <sup>-</sup>	1-	thiocyanate ion
Ni <sup>2+</sup>	2+	nickel(II) ion or nickelous ion	O <sup>2-</sup>	2-	oxide ion
Pb <sup>2+</sup>	2+	lead(II) ion or plumbous ion	S <sup>2-</sup>	2-	sulfide ion
Sn <sup>2+</sup>	2+	tin(II) ion or stannous ion	HSO <sub>3</sub> <sup>-</sup>	1-	hydrogen sulfite ion or bisulfite ion
Fe <sup>2+</sup>	2+	iron(II) ion or ferrous ion	SO <sub>3</sub> <sup>2-</sup>	2-	sulfite ion
Fe <sup>3+</sup>	3+	iron(III) ion or ferric ion	HSO <sub>4</sub> <sup>-</sup>	1-	hydrogen sulfate ion or bisulfate ion
Al <sup>3+</sup>	3+	aluminum ion	SO <sub>4</sub> <sup>2-</sup>	2-	sulfate ion
Cr <sup>3+</sup>	3+	chromium(III) ion or chromic ion	HCO <sub>3</sub> <sup>-</sup>	1-	hydrogen carbonate ion or bicarbonate ion
			CO <sub>3</sub> <sup>2-</sup>	2-	carbonate ion
			CrO <sub>4</sub> <sup>2-</sup>	2-	chromate ion
			Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup>	2-	dichromate ion
			PO <sub>4</sub> <sup>3-</sup>	3-	phosphate ion
			AsO <sub>4</sub> <sup>3-</sup>	3-	arsenate ion

If you don't already know them, you should learn these common prefixes.

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

Although many nonmetals can exhibit different oxidation numbers, their oxidation numbers are *not* properly indicated by Roman numerals or suffixes. Instead, elemental proportions in binary covalent compounds are indicated by using a *prefix* system for both elements. The Greek and Latin prefixes for one through ten are *mono-*, *di-*, *tri-*, *tetra-*, *penta-*, *hexa-*, *hepta-*, *octa-*, *nona-*, and *deca-*. The prefix “mono-” is omitted for both elements except in the common name for CO, carbon monoxide. We use the minimum number of prefixes needed to name a compound unambiguously. The final “a” in a prefix is omitted when the nonmetal stem begins with the letter “o”; we write “heptoxide,” not “heptaoxide.”

Formula	Name	Formula	Name
SO <sub>2</sub>	sulfur dioxide	Cl <sub>2</sub> O <sub>7</sub>	dichlorine heptoxide
SO <sub>3</sub>	sulfur trioxide	CS <sub>2</sub>	carbon disulfide
N <sub>2</sub> O <sub>4</sub>	dinitrogen tetroxide	SF <sub>4</sub>	sulfur tetrafluoride
As <sub>4</sub> O <sub>6</sub>	tetraarsenic hexoxide	SF <sub>6</sub>	sulfur hexafluoride

**Binary acids** are compounds in which H is bonded to a Group VIA element other than O or to a Group VIIA element; they act as acids when dissolved in water. *The pure compounds are named as typical binary compounds.* Their aqueous solutions are named by modifying the characteristic stem of the nonmetal with the prefix “hydro-” and the suffix “-ic” followed by the word “acid.” The stem for sulfur in this instance is “sulfur” rather than “sulf.”

Formula	Name of Compound	Name of Aqueous Solution
HCl	hydrogen chloride	hydrochloric acid, HCl(aq)
HF	hydrogen fluoride	hydrofluoric acid, HF(aq)
H <sub>2</sub> S	hydrogen sulfide	hydrosulfuric acid, H <sub>2</sub> S(aq)
HCN	hydrogen cyanide	hydrocyanic acid, HCN(aq)

In later chapters we will learn additional systematic rules for naming more complex compounds.

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## 4-6 NAMING TERNARY ACIDS AND THEIR SALTS

A ternary compound consists of three elements. **Ternary acids (oxoacids)** are compounds of hydrogen, oxygen, and (usually) a nonmetal. Nonmetals that exhibit more than one oxidation state form more than one ternary acid. These ternary acids differ in the number of oxygen atoms they contain. The suffixes “-ous” and “-ic” following the stem name of the central element indicate lower and higher oxidation states, respectively. One common ternary acid of each nonmetal is (somewhat arbitrarily) designated as the “-ic” acid. That is, it is named by putting the element stem before the “-ic” suffix. The common ternary “-ic acids” are shown in Table 4-12. It is important to learn the names and formulas of these acids because the names of all other ternary acids and salts are derived from them. There are no common “-ic” ternary acids for the omitted nonmetals.

The oxoacid with the central element in the highest oxidation state usually contains more O atoms. Oxoacids with their central elements in lower oxidation states usually have fewer O atoms.

**TABLE 4-12** Formulas of Some “-ic” Acids

Periodic Group of Central Elements				
IIA	IVA	VA	VIA	VIIA
<sup>+3</sup> H <sub>3</sub> BO <sub>3</sub> boric acid	<sup>+4</sup> H <sub>2</sub> CO <sub>3</sub> carbonic acid	<sup>+5</sup> HNO <sub>3</sub> nitric acid		
	<sup>+4</sup> H <sub>4</sub> SiO <sub>4</sub> silicic acid	<sup>+5</sup> H <sub>3</sub> PO <sub>4</sub> phosphoric acid	<sup>+6</sup> H <sub>2</sub> SO <sub>4</sub> sulfuric acid	<sup>+5</sup> HClO <sub>3</sub> chloric acid
		<sup>+5</sup> H <sub>3</sub> AsO <sub>4</sub> arsenic acid	<sup>+6</sup> H <sub>2</sub> SeO <sub>4</sub> selenic acid	<sup>+5</sup> HBrO <sub>3</sub> bromic acid
			<sup>+6</sup> H <sub>6</sub> TeO <sub>6</sub> telluric acid	<sup>+5</sup> HIO <sub>3</sub> iodic acid

Note that the oxidation state of the central atom is equal to its periodic group number, except for the halogens.

Acids containing *one fewer oxygen atom* per central atom are named in the same way except that the “-ic” suffix is changed to “-ous.” The oxidation number of the central element is *lower by 2* in the “-ous” acid than in the “-ic” acid.

Formula	Ox. No.	Name	Formula	Ox. No.	Name
H <sub>2</sub> SO <sub>3</sub>	+4	sulfurous acid	H <sub>2</sub> SO <sub>4</sub>	+6	sulfuric acid
HNO <sub>2</sub>	+3	nitrous acid	HNO <sub>3</sub>	+5	nitric acid
H <sub>2</sub> SeO <sub>3</sub>	+4	selenous acid	H <sub>2</sub> SeO <sub>4</sub>	+6	selenic acid
HBrO <sub>2</sub>	+3	bromous acid	HBrO <sub>3</sub>	+5	bromic acid

Ternary acids that have one fewer O atom than the “-ous” acids (two fewer O atoms than the “-ic” acids) are named using the prefix “hypo-” and the suffix “-ous.” These are acids in which the oxidation state of the central nonmetal is *lower by 2* than that of the central nonmetal in the “-ous acids.”

Formula	Ox. No.	Name
HClO	+1	<i>hypochlorous acid</i>
H <sub>3</sub> PO <sub>2</sub>	+1	<i>hypophosphorous acid</i>
HIO	+1	<i>hypoiodous acid</i>
H <sub>2</sub> N <sub>2</sub> O <sub>2</sub>	+1	<i>hyponitrous acid</i>

Notice that H<sub>2</sub>N<sub>2</sub>O<sub>2</sub> has a 1:1 ratio of nitrogen to oxygen, as would the hypothetical HNO.

Acids containing *one more oxygen atom* per central nonmetal atom than the normal “-ic acid” are named “*perstemic*” acids.

Formula	Ox. No.	Name
HClO <sub>4</sub>	+7	<i>perchloric acid</i>
HBrO <sub>4</sub>	+7	<i>perbromic acid</i>
HIO <sub>4</sub>	+7	<i>periodic acid</i>

The oxoacids of chlorine can be summarized as follows:

Formula	Ox. No.	Name
HClO	+1	<i>hypochlorous acid</i>
HClO <sub>2</sub>	+3	<i>chlorous acid</i>
HClO <sub>3</sub>	+5	<i>chloric acid</i>
HClO <sub>4</sub>	+7	<i>perchloric acid</i>

**Ternary salts** are compounds that result from replacing the hydrogen in a ternary acid with another ion. They usually contain metal cations or the ammonium ion. As with binary compounds, the cation is named first. The name of the anion is based on the name of the ternary acid from which it is derived.


An anion derived from a ternary acid with an “-ic” ending is named by dropping the “-ic acid” and replacing it with “-ate.” An anion derived from an “-ous acid” is named by replacing the suffix “-ous acid” with “-ite.” The “per-” and “hypo-” prefixes are retained.

Formula	Name
$(\text{NH}_4)_2\text{SO}_4$	ammonium sulfate ( $\text{SO}_4^{2-}$ , from sulfuric acid, $\text{H}_2\text{SO}_4$ )
$\text{KNO}_3$	potassium nitrate ( $\text{NO}_3^-$ , from nitric acid, $\text{HNO}_3$ )
$\text{Ca}(\text{NO}_2)_2$	calcium nitrite ( $\text{NO}_2^-$ , from nitrous acid, $\text{HNO}_2$ )
$\text{LiClO}_4$	lithium perchlorate ( $\text{ClO}_4^-$ , from perchloric acid, $\text{HClO}_4$ )
$\text{FePO}_4$	iron(III) phosphate ( $\text{PO}_4^{3-}$ , from phosphoric acid, $\text{H}_3\text{PO}_4$ )
$\text{NaClO}$	sodium hypochlorite ( $\text{ClO}^-$ , from hypochlorous acid, $\text{HClO}$ )

Acidic salts contain anions derived from ternary polyprotic acids in which one or more acidic hydrogen atoms remain. These salts are named as if they were the usual type of ternary salt, with the word “hydrogen” or “dihydrogen” inserted after the name of the cation to show the number of acidic hydrogen atoms.

Formula	Name	Formula	Name
$\text{NaHSO}_4$	sodium hydrogen sulfate	$\text{KH}_2\text{PO}_4$	potassium dihydrogen phosphate
$\text{NaHSO}_3$	sodium hydrogen sulfite	$\text{K}_2\text{HPO}_4$	potassium hydrogen phosphate
		$\text{NaHCO}_3$	sodium hydrogen carbonate

An older, commonly used method (which is not recommended by the IUPAC, but which is widely used in commerce) involves the use of the prefix “bi-” attached to the name of the anion to indicate the presence of an acidic hydrogen. According to this system,  $\text{NaHSO}_4$  is called sodium bisulfate and  $\text{NaHCO}_3$  is called sodium bicarbonate.

 **Problem-Solving Tip:** *Naming Ternary Acids and Their Anions*

The following table might help you to remember the names of the ternary acids and their ions. First learn the formulas of the acids mentioned earlier that end with “-ic acid.” Then relate possible other acids to the following table. The stem (XXX) represents the stem of the name, for example, “nitr,” “sulfur,” or “chlor.”

Decreasing oxidation number of central atom	Ternary Acid		Anion	Decreasing number of oxygen atoms on central atom
	<i>perXXXic acid</i>		<i>perXXXate</i>	
	<i>XXXic acid</i>		<i>XXXate</i>	
	<i>XXXous acid</i>		<i>XXXite</i>	
	<i>hypoXXXous acid</i>		<i>hypoXXXite</i>	

## CLASSIFYING CHEMICAL REACTIONS

We now discuss chemical reactions in further detail. We classify them as oxidation–reduction reactions, combination reactions, decomposition reactions, displacement reactions, and metathesis reactions. The last type can be further described as precipitation reactions and acid–base (neutralization) reactions. We will see that many reactions, especially oxidation–reduction reactions, fit into more than one category, and that some reactions do not fit neatly into any of them. As we study different kinds of chemical reactions, we will learn to predict the products of other similar reactions. In Chapter 6 we will describe typical



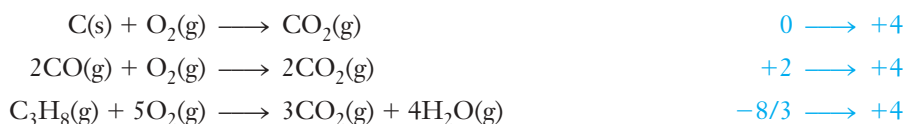
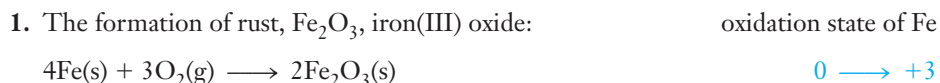
CD-ROM Screen 5.6, Types of Reactions in Aqueous Solution



reactions of hydrogen, oxygen, and their compounds. These reactions will illustrate periodic relationships with respect to chemical properties. It should be emphasized that our system is not an attempt to transform nature so that it fits into small categories but rather an effort to give some order to our many observations of nature.

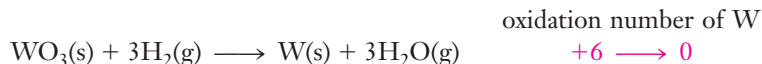
#### 4-7 OXIDATION–REDUCTION REACTIONS: INTRODUCTION

The term “oxidation” originally referred to the combination of a substance with oxygen. This results in an increase in the oxidation number of an element in that substance. According to the original definition, the following reactions involve oxidation of the substance shown on the far left of each equation. Oxidation numbers are shown for *one* atom of the indicated kind.



Oxidation number is a formal concept adopted for our convenience. The numbers are determined by relying on rules. These rules can result in a fractional oxidation number, as shown here. This does not mean that electronic charges are split.

Originally, *reduction* described the removal of oxygen from a compound. Oxide ores are reduced to metals (a very real reduction in mass). For example, tungsten for use in light bulb filaments can be prepared by reduction of tungsten(VI) oxide with hydrogen at 1200°C:



The terms “oxidation number” and “oxidation state” are used interchangeably.

Tungsten is reduced, and its oxidation state decreases from +6 to zero. Hydrogen is oxidized from zero to the +1 oxidation state. The terms “oxidation” and “reduction” are now applied much more broadly.

In biological systems, *reduction* often corresponds to the addition of hydrogen to molecules or polyatomic ions and *oxidation* often corresponds to the removal of hydrogen.

**Oxidation** is an increase in oxidation number and corresponds to the loss, or apparent loss, of electrons. **Reduction** is a decrease in oxidation number and corresponds to a gain, or apparent gain, of electrons.

Electrons are neither created nor destroyed in chemical reactions. So oxidation and reduction always occur simultaneously, and to the same extent, in ordinary chemical reactions. In the four equations cited previously as *examples of oxidation*, the oxidation numbers of iron and carbon atoms increase as they are oxidized. In each case oxygen is reduced as its oxidation number decreases from zero to  $-2$ .

Because oxidation and reduction occur simultaneously in all of these reactions, they are referred to as oxidation–reduction reactions. For brevity, we usually call them **redox** reactions. Redox reactions occur in nearly every area of chemistry and biochemistry. We need to be able to identify oxidizing agents and reducing agents and to balance oxidation–reduction equations. These skills are necessary for the study of electrochemistry in Chapter 21. Electrochemistry involves electron transfer between physically separated oxidizing and



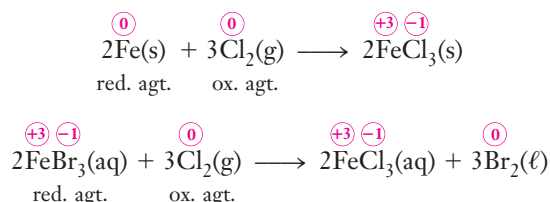
reducing agents and interconversions between chemical energy and electrical energy. These skills are also fundamental to the study of biology, biochemistry, environmental science, and materials science.

SCILINKS

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**Oxidizing agents** are species that (1) oxidize other substances, (2) contain atoms that are reduced, and (3) gain (or appear to gain) electrons. **Reducing agents** are species that (1) reduce other substances, (2) contain atoms that are oxidized, and (3) lose (or appear to lose) electrons.

The following equations represent examples of redox reactions. Oxidation numbers are shown above the formulas, and oxidizing and reducing agents are indicated:



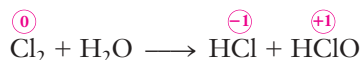
Equations for redox reactions can also be written as total ionic and net ionic equations. For example, the second equation may also be written as:



We distinguish between oxidation numbers and actual charges on ions by denoting oxidation numbers as  $+n$  or  $-n$  in red circles just above the symbols of the elements, and actual charges as  $n+$  or  $n-$  above and to the right of formulas of ions. The spectator ions,  $\text{Fe}^{3+}$ , do not participate in electron transfer. Their cancellation allows us to focus on the oxidizing agent,  $\text{Cl}_2(\text{g})$ , and the reducing agent,  $\text{Br}^{-}(\text{aq})$ .

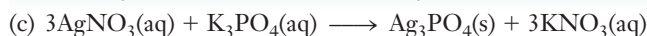
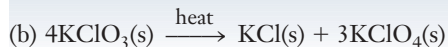


A **disproportionation reaction** is a redox reaction in which the same element is oxidized and reduced. An example is:



### EXAMPLE 4-5 Redox Reactions

Write each of the following formula unit equations as a net ionic equation if the two differ. Which ones are redox reactions? For the redox reactions, identify the oxidizing agent, the reducing agent, the species oxidized, and the species reduced.



#### Plan

To write ionic equations, we must recognize compounds that are (1) soluble in water and (2) ionized or dissociated in aqueous solutions. To determine which are oxidation–reduction reactions, we should assign an oxidation number to each element.

The following abbreviations are widely used:

ox. no. = oxidation number

ox. agt. = oxidizing agent

red. agt. = reducing agent



Charles Steele

Iron reacting with chlorine to form iron(III) chloride.

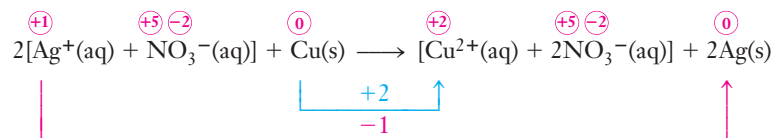


James W. Morganthaler

Metallic silver formed by immersing a spiral of copper wire in a silver nitrate solution (see Example 4-5a).

**Solution**

(a) According to the solubility guidelines (Section 4-2 Part 5), both silver nitrate,  $\text{AgNO}_3$ , and copper(II) nitrate,  $\text{Cu}(\text{NO}_3)_2$ , are water-soluble ionic compounds. The total ionic equation and oxidation numbers are

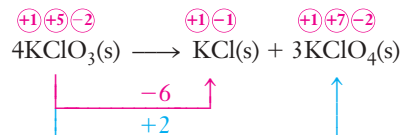


The nitrate ions,  $\text{NO}_3^-$ , are spectator ions. Canceling them from both sides gives the net ionic equation:



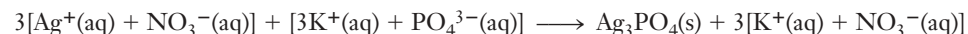
This is a redox equation. The oxidation number of silver decreases from +1 to zero; silver ion is reduced and is the oxidizing agent. The oxidation number of copper increases from zero to +2; copper is oxidized and is the reducing agent.

(b) This reaction involves only solids, so there are no ions in solution, and the formula unit and net ionic equations are identical. It is a redox reaction:

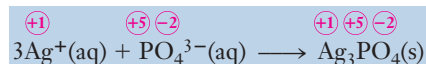


Chlorine is reduced from +5 in  $\text{KClO}_3$  to the  $-1$  oxidation state in  $\text{KCl}$ ; the oxidizing agent is  $\text{KClO}_3$ . Chlorine is oxidized from +5 in  $\text{KClO}_3$  to the +7 oxidation state in  $\text{KClO}_4$ .  $\text{KClO}_3$  is also the reducing agent. This is a disproportionation reaction. We see that  $\text{KClO}_3$  is both the oxidizing agent and the reducing agent.

(c) The solubility guidelines indicate that all these salts are soluble except for silver phosphate,  $\text{Ag}_3\text{PO}_4$ . The total ionic equation is

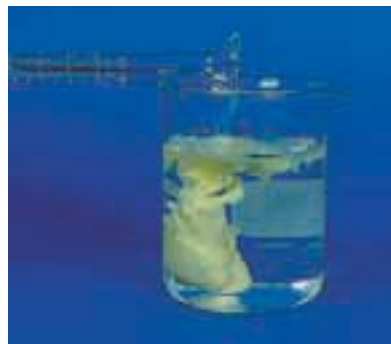


Eliminating the spectator ions gives the net ionic equation:



There are no changes in oxidation numbers; this is not a redox reaction.

You should now work Exercises 66 and 69.



James W. Mongenthaler

The reaction of  $\text{AgNO}_3(\text{aq})$  and  $\text{K}_3\text{PO}_4(\text{aq})$  is a precipitation reaction (see Example 4-5c).

**Problem-Solving Tip:** A Foolproof Way to Recognize a Redox Reaction

You can always recognize a redox reaction by analyzing oxidation numbers. First determine the oxidation number of each element wherever it appears in the reaction. If no elements change in oxidation numbers, the reaction is not an oxidation–reduction reaction. If changes do occur, the reaction is an oxidation–reduction reaction. Remember that oxidation and reduction must always occur together; if some atoms increase in oxidation numbers, then others must decrease.

In Chapter 11 we will learn to balance redox equations and to carry out stoichiometric calculations using the balanced equations.

## 4-8 COMBINATION REACTIONS

Reactions in which two or more substances combine to form a compound are called **combination reactions**.

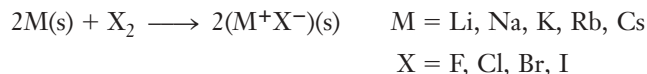
They may involve (1) the combination of two elements to form a compound, (2) the combination of an element and a compound to form a single new compound, or (3) the combination of two compounds to form a single new compound. Let's examine some of these reactions.

### 1 Element + Element → Compound

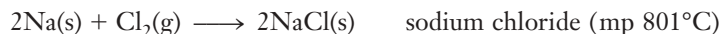
For this type of combination reaction, each element goes from an uncombined state, where its oxidation state is zero, to a combined state in a compound, where its oxidation state is not zero. Thus reactions of this type are oxidation–reduction reactions (see Section 4-7).

#### *Metal + Nonmetal → Binary Ionic Compound*

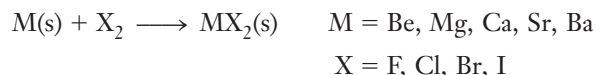
Most metals react with most nonmetals to form binary ionic compounds. The Group IA metals combine with the Group VIIA nonmetals to form binary *ionic* compounds with the general formula MX (Section 7-2):



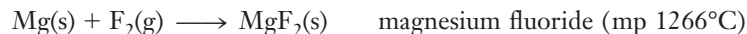
This general equation thus represents the 20 combination reactions that form the ionic compounds listed in Table 4-13. Sodium, a silvery-white metal, combines with chlorine, a pale green gas, to form sodium chloride, or ordinary table salt. All members of both families undergo similar reactions.



As we might expect, the Group IIA metals also combine with the Group VIIA nonmetals to form binary compounds. Except for  $\text{BeCl}_2$ ,  $\text{BeBr}_2$ , and  $\text{BeI}_2$ , these are ionic compounds. In general terms these combination reactions may be represented as:



Consider the reaction of magnesium with fluorine to form magnesium fluoride:



Because all the IIA and VIIA elements undergo similar reactions, the general equation, written above, represents 20 reactions. We omit radium and astatine, the rare and highly radioactive members of the families.

#### *Nonmetal + Nonmetal → Binary Covalent Compound*

When two nonmetals combine with each other, they form binary *covalent* compounds. In such reactions, the oxidation number of the element with the more positive oxidation



Charles D. Winters

Potassium, a metal, reacts with chlorine, a nonmetal, to form potassium chloride, KCl. The reaction releases energy in the form of heat and light.

Another important reaction of this kind is the formation of metal oxides (Section 6-8).

**TABLE 4-13** Alkali Metal Halides: Compounds Formed by Group IA and VIIA Elements

LiF	LiCl	LiBr	LiI
NaF	NaCl	NaBr	NaI
KF	KCl	KBr	KI
RbF	RbCl	RbBr	RbI
CsF	CsCl	CsBr	CsI

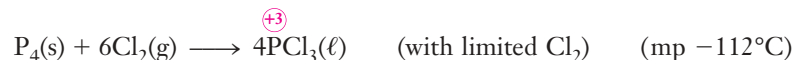


Charles D. Winters

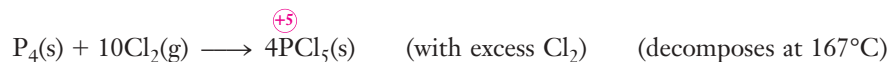
Phosphorus and chlorine, two nonmetals, react to form phosphorus pentachloride,  $\text{PCl}_5$ .

Nonmetals in odd-numbered periodic groups favor odd oxidation numbers, whereas those in even-numbered groups favor even oxidation numbers in their compounds. The *maximum* oxidation number for a representative element is equal to its periodic group number. For example, sulfur (Group VIA) can form both  $\text{SF}_4$  and  $\text{SF}_6$ .

number is often variable, depending on reaction conditions. For example, phosphorus (Group VA) combines with a *limited amount* of chlorine to form phosphorus trichloride, in which phosphorus exhibits the +3 oxidation state.



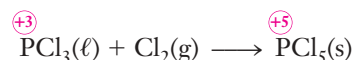
With an excess of chlorine, the product is phosphorus pentachloride, which contains phosphorus in the +5 oxidation state:



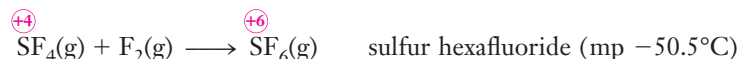
In general, *a higher oxidation state of a nonmetal is formed when it reacts with an excess of another nonmetal*. There are many more reactions in which two elements combine to form a compound (see Sections 6-7 and 6-8).

## 2 Compound + Element $\rightarrow$ Compound

Phosphorus in the +3 oxidation state in  $\text{PCl}_3$  molecules can be converted to the +5 state in  $\text{PCl}_5$  by combination with chlorine:



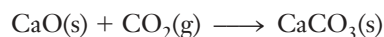
Likewise, sulfur in the +4 state is converted to the +6 state when  $\text{SF}_4$  reacts with fluorine to form  $\text{SF}_6$ :



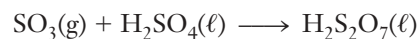
Combination reactions of this type are also oxidation–reduction reactions.

## 3 Compound + Compound $\rightarrow$ Compound

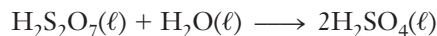
An example of reactions in this category is the combination of calcium oxide with carbon dioxide to produce calcium carbonate:



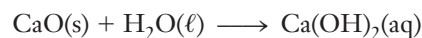
Pyrosulfuric acid is produced by dissolving sulfur trioxide in concentrated sulfuric acid:



Pyrosulfuric acid,  $\text{H}_2\text{S}_2\text{O}_7$ , is then diluted with water to make  $\text{H}_2\text{SO}_4$ :



Oxides of the Group IA and IIA metals react with water to form metal hydroxides, e.g.:



## 4-9 DECOMPOSITION REACTIONS

Decomposition reactions can be considered as the opposite of combination reactions.

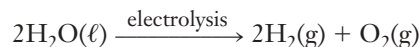
**Decomposition reactions** are those in which a compound decomposes to produce (1) two elements, (2) one or more elements *and* one or more compounds, or (3) two or more compounds.



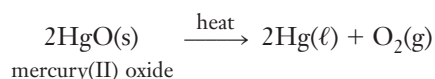
Examples of each type follow.

### 1 Compound → Element + Element

The electrolysis of water produces two elements by the decomposition of a compound. A compound that ionizes, such as  $\text{H}_2\text{SO}_4$ , is added to increase the conductivity of water and the rate of the reaction (Figure 1-8), but it does not participate in the reaction:

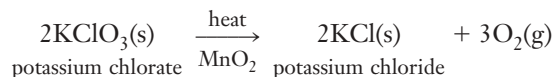


Small amounts of oxygen can be prepared by the thermal decomposition of certain oxygen-containing compounds. Some metal oxides, such as mercury(II) oxide,  $\text{HgO}$ , decompose on heating to produce oxygen:



### 2 Compound → Compound + Element

The alkali metal chlorates, such as  $\text{KClO}_3$ , decompose when heated to produce the corresponding chlorides and liberate oxygen. Potassium chlorate is a common laboratory source of small amounts of oxygen:



Nitrate salts of alkali metals or alkaline earth metals decompose to form metal nitrites and oxygen gas.

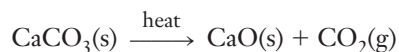


Hydrogen peroxide decomposes to form water and oxygen.



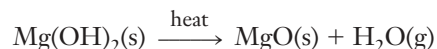
### 3 Compound → Compound + Compound

The thermal decomposition of calcium carbonate (limestone) and other carbonates produces two compounds, a metal oxide and carbon dioxide:



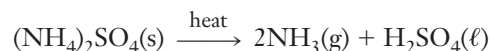
This is an important reaction in the production of cement. Calcium oxide is also used as a base in industrial processes.

When some solid hydroxides are heated, they decompose to form a metal oxide and water vapor.



Magnesium oxide,  $\text{MgO}$ , is pressed into sheets for use as a thermal insulating material in oven walls.

Ammonium salts lose ammonia.



Charles D. Winters

Mercury(II) oxide, a red compound, decomposes when heated into the two elements: mercury (a metal) and oxygen (a nonmetal). Mercury vapor condenses on the cooler upper portion of the test tube.

Manganese dioxide,  $\text{MnO}_2$ , is used as a catalyst, a substance that speeds up a chemical reaction but is not consumed. Here it allows the decomposition to occur at a lower temperature.

Alkali metal carbonates do not decompose when heated.

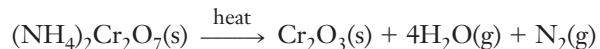
A decomposition reaction may or may not also be an oxidation–reduction reaction. You can always identify a redox reaction by determining the oxidation state of each element in each occurrence in the reaction (see the Problem-Solving Tip in Section 4-7).



Charles D. Winters

Solid ammonium dichromate,  $[(\text{NH}_4)_2\text{Cr}_2\text{O}_7]$ , orange decomposes when heated into chromium(III) oxide,  $(\text{Cr}_2\text{O}_3)$ , green, nitrogen, and steam (water vapor). This reaction is sometimes demonstrated as the “classroom volcano,” but it must be done with extreme caution due to the carcinogenic (cancer-causing) nature of  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ .

If the ammonium salt contains an anion that is a strong oxidizing agent (e.g., nitrate, nitrite, or dichromate), its decomposition reaction produces an oxide, water (as vapor at high temperatures), and nitrogen gas. Such a reaction is a redox reaction.



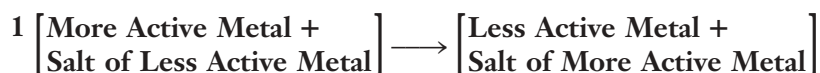
#### 4-10 DISPLACEMENT REACTIONS

Reactions in which one element displaces another from a compound are called **displacement reactions**.

These reactions are always redox reactions. The more readily a metal undergoes oxidation, the more active we say it is.

Active metals displace less active metals or hydrogen from their compounds in aqueous solution to form the oxidized form of the more active metal and the reduced (free metal) form of the other metal or hydrogen.

In Table 4-14, the most active metals are listed at the top of the first column. These metals tend to react to form their oxidized forms (cations). Elements at the bottom of the activity series (the first column of Table 4-14) tend to remain in their reduced form. They are easily converted from their oxidized forms to their reduced forms.



The reaction of copper with silver nitrate that was described in detail in Section 4-3 is typical. Please refer to it.

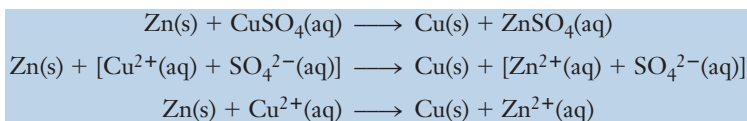
#### EXAMPLE 4-6 Displacement Reaction

A large piece of zinc metal is placed in a copper(II) sulfate,  $\text{CuSO}_4$ , solution. The blue solution becomes colorless as copper metal falls to the bottom of the container. The resulting solution contains zinc sulfate,  $\text{ZnSO}_4$ . Write balanced formula unit, total ionic, and net ionic equations for the reaction.

##### Plan

The metals zinc and copper are *not* ionized or dissociated in contact with  $\text{H}_2\text{O}$ . Both  $\text{CuSO}_4$  and  $\text{ZnSO}_4$  are soluble salts (solubility guideline 5), so they are written in ionic form.

##### Solution



In this *displacement reaction*, the more active metal, zinc, displaces the ions of the less active metal, copper, from aqueous solution.

*You should now work Exercise 78.*



James W. Morganthaler

A strip of zinc metal was placed in a blue solution of copper(II) sulfate,  $\text{CuSO}_4$ . The copper has been displaced from solution and has fallen to the bottom of the beaker. The resulting zinc sulfate,  $\text{ZnSO}_4$ , solution is colorless.

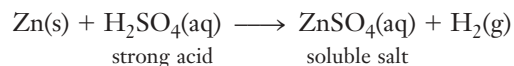
**TABLE 4-14** Activity Series of Some Elements

Element	Common Reduced Form	Common Oxidized Forms
Li	Li	Li <sup>+</sup>
K	K	K <sup>+</sup>
Ca	Ca	Ca <sup>2+</sup>
Na	Na	Na <sup>+</sup>
Mg	Mg	Mg <sup>2+</sup>
Al	Al	Al <sup>3+</sup>
Mn	Mn	Mn <sup>2+</sup>
Zn	Zn	Zn <sup>2+</sup>
Cr	Cr	Cr <sup>3+</sup> , Cr <sup>6+</sup>
Fe	Fe	Fe <sup>2+</sup> , Fe <sup>3+</sup>
Cd	Cd	Cd <sup>2+</sup>
Co	Co	Co <sup>2+</sup>
Ni	Ni	Ni <sup>2+</sup>
Sn	Sn	Sn <sup>2+</sup> , Sn <sup>4+</sup>
Pb	Pb	Pb <sup>2+</sup> , Pb <sup>4+</sup>
H (a nonmetal)	H <sub>2</sub>	H <sup>+</sup>
Sb (a metalloid)	Sb	Sb <sup>3+</sup>
Cu	Cu	Cu <sup>+</sup> , Cu <sup>2+</sup>
Hg	Hg	Hg <sub>2</sub> <sup>2+</sup> , Hg <sup>2+</sup>
Ag	Ag	Ag <sup>+</sup>
Pt	Pt	Pt <sup>2+</sup> , Pt <sup>4+</sup>
Au	Au	Au <sup>+</sup> , Au <sup>3+</sup>

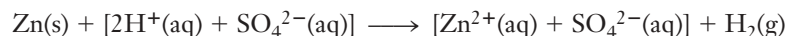
**TOPIC:** Activity Series  
**GO TO:** [www.scilinks.org](http://www.scilinks.org)  
**sciLINKS CODE:** WCH0470

## 2 [Active Metal + Nonoxidizing Acid] → [Hydrogen + Salt of Acid]

A common method for preparing small amounts of hydrogen involves the reaction of active metals with nonoxidizing acids, such as HCl and H<sub>2</sub>SO<sub>4</sub>. For example, when zinc is dissolved in H<sub>2</sub>SO<sub>4</sub>, the reaction produces zinc sulfate; hydrogen is displaced from the acid, and it bubbles off as gaseous H<sub>2</sub>. The formula unit equation for this reaction is



Both sulfuric acid (in very dilute solution) and zinc sulfate exist primarily as ions, so the total ionic equation is



Elimination of spectator ions from the total ionic equation gives the net ionic equation:

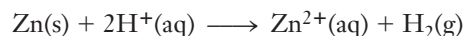


Table 4-14 lists the **activity series**. When any metal listed above hydrogen in this series is added to a solution of a *nonoxidizing* acid such as hydrochloric acid, HCl, and sulfuric acid, H<sub>2</sub>SO<sub>4</sub>, the metal dissolves to produce hydrogen, and a salt is formed. HNO<sub>3</sub> is the common *oxidizing acid*. It reacts with active metals to produce oxides of nitrogen, but *not* hydrogen, H<sub>2</sub>.

H<sub>2</sub>SO<sub>4</sub> can function as an oxidizing agent with other substances, but it is not an oxidizing agent in its reaction with active metals.



Charles Steele

Zinc reacts with dilute H<sub>2</sub>SO<sub>4</sub> to produce H<sub>2</sub> and a solution that contains ZnSO<sub>4</sub>. This is a displacement reaction.



## Troublesome Displacement Reactions

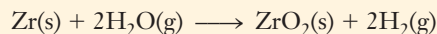
The deterioration of the Statue of Liberty and the damage done at the Three Mile Island and Chernobyl nuclear facilities are just a few of the major problems that have resulted from ignorance about chemical reactivity.

When originally constructed over one hundred years ago the Statue of Liberty had a 200,000-pound outer copper skin supported by a framework of 2000 iron bars. First, oxygen in the air oxidized the copper skin to form copper oxide. In a series of reactions, iron (the more active metal) then reduced the  $\text{Cu}^{2+}$  ions in copper oxide.



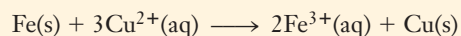
Over the years, the supporting iron frame was reduced to less than half its original thickness; this made necessary the repairs done to the statue before the celebration of its 100th birthday on July 4, 1986.

Two major nuclear power plant accidents, one at Three Mile Island near Harrisburg, Pennsylvania, in 1979 and the other at Chernobyl in Ukraine in 1986, were also unexpected consequences of chemical reactivity. In each case, cooling pump failures sent temperatures soaring above  $340^\circ\text{C}$ . Like aluminum, zirconium (used in building the reactors) forms an oxide coating that shields it from further reactions. However, that protective coating breaks down at high temperatures. Without its protective coating, zirconium reacts with steam.



At Three Mile Island, this displacement reaction produced a 1000-cubic foot bubble of hydrogen gas. Because hydrogen is easily ignited by a spark, the nuclear power plant was in real danger of a complete meltdown until the hydrogen could be removed.

During the Middle Ages ( $\sim\text{AD } 400\text{--}1400$ ), another displacement reaction completely misled alchemists into foolishly pursuing a philosophers' stone that was believed to have the power to turn base metals such as iron and lead into more precious metals such as silver and gold. The alchemists' ignorance of relative activities of metals led them to believe that they had turned iron into a more precious metal when they inserted an iron rod into a blue copper(II) sulfate solution. In fact, the following displacement reaction had occurred, plating shiny copper metal onto the iron rod:

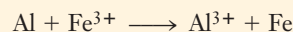


In the 1960s and 1970s, some automobile manufacturers showed their ignorance of chemical reactivity by building cars with aluminum water pumps and aluminum engine heads



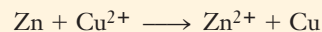
Patricia Caulfield/Photo Researchers, Inc.

attached to cast-iron engine blocks. These water pumps often leaked and the engine heads quickly deteriorated. These problems occurred as the more active aluminum reacted with iron(II) oxide (formed when the iron engine reacted with atmospheric oxygen).



Some dentists have made similar mistakes by placing gold caps over teeth that are adjacent to existing fillings. The slightly oxidized gold can react with a dental amalgam filling (an alloy of silver, tin, copper, and mercury). As the dental amalgam is oxidized, it dissolves in saliva to produce a persistent metallic taste in the patient's mouth.

When plumbers connect galvanized pipes (iron pipes coated with zinc) to copper pipes, copper ions oxidize the zinc coating and expose the underlying iron, allowing it to rust. The displacement reaction that occurs is



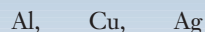
Once the zinc coating has been punctured on an iron pipe, oxidation of the iron pipes occurs rapidly because iron is a more active metal than copper.

It is important to keep in mind that a variety of reactions other than the displacement reactions discussed here probably take place. For example, less active metals (such as copper) can conduct electrons from the metals being oxidized to oxidizing agents (such as oxygen or the oxide of nitrogen and sulfur) that are present in the atmosphere. Oxygen plays an important role in all these displacement examples.

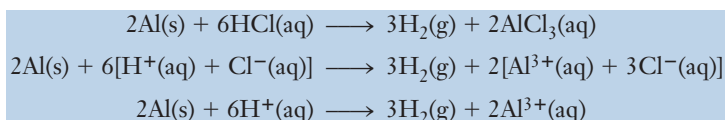
*Ronald DeLorenzo  
Middle Georgia College*

**EXAMPLE 4-7 Displacement Reaction**

Which of the following metals can displace hydrogen from hydrochloric acid solution? Write balanced formula unit, total ionic, and net ionic equations for reactions that can occur.

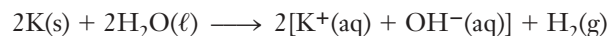
**Plan**

The activity series of the metals, Table 4-14, tells us that copper and silver *do not* displace hydrogen from solutions of nonoxidizing acids. Aluminum is an active metal that can displace  $\text{H}_2$  from HCl and form aluminum chloride.

**Solution**

You should now work Exercises 77 and 79.

Very active metals can even displace hydrogen from water. The reaction of potassium, or another metal of Group IA, with water is also a *displacement reaction*:



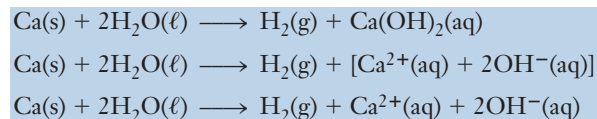
Such reactions of very active metals of Group IA are dangerous, however, because they generate enough heat to cause explosive ignition of the hydrogen (Figure 4-4).

**EXAMPLE 4-8 Displacement Reaction**

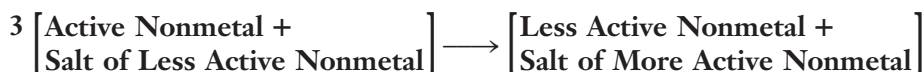
Which of the following metals can displace hydrogen from water at room temperature? Write balanced formula unit, total ionic, and net ionic equations for reactions that can occur.

**Plan**

The activity series, Table 4-14, tells us that tin and mercury *cannot* displace hydrogen from water. Calcium is a very active metal (see Table 4-14) that displaces hydrogen from cold water and forms calcium hydroxide, a strong base.

**Solution**

You should now work Exercise 85.



Many *nonmetals* displace less active nonmetals from combination with a metal or other cation. For example, when chlorine is bubbled through a solution containing bromide ions



Charles Steele

Aluminum displaces  $\text{H}_2$  from a hydrochloric acid solution.



Charles Steele

**Figure 4-4** Potassium, like other Group IA metals, reacts vigorously with water. For this photograph, the room was completely dark, and all the light you see here was produced by dropping a small piece of potassium into a beaker of water.



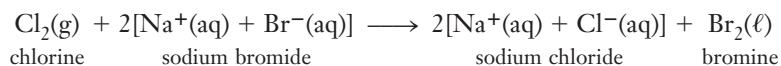


Charles D. Winters

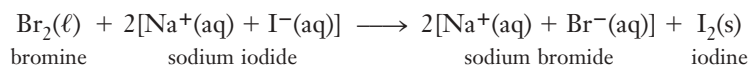
The displacement reaction of calcium with water at room temperature produces bubbles of hydrogen.

Activity of the halogens decreases going down the group in the periodic table.

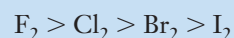
(derived from a soluble ionic salt such as sodium bromide, NaBr), chlorine displaces bromide ions to form elemental bromine and chloride ions (as aqueous sodium chloride):



Similarly, when bromine is added to a solution containing iodide ions, the iodide ions are displaced by bromine to form iodine and bromide ions:



*Each halogen will displace less active (heavier) halogens from their binary salts; that is, the order of decreasing activities is*



Conversely, a halogen will *not* displace more active (lighter) members from their salts:



### EXAMPLE 4-9 Displacement Reactions

Which of the following combinations would result in a displacement reaction? Write balanced formula unit, total ionic, and net ionic equations for reactions that occur.

- (a)  $\text{I}_2(\text{s}) + \text{NaBr}(\text{aq}) \longrightarrow$   
 (b)  $\text{Cl}_2(\text{g}) + \text{NaI}(\text{aq}) \longrightarrow$   
 (c)  $\text{Br}_2(\ell) + \text{NaCl}(\text{aq}) \longrightarrow$



(a)



(b)

(a) Bromine,  $\text{Br}_2$ , in water (pale orange) is poured into an aqueous solution of NaI, the top layer in the cylinder. (b)  $\text{Br}_2$  displaces  $\text{I}^-$  from solution and forms solid iodine,  $\text{I}_2$ . The  $\text{I}_2$  dissolves in water to give a brown solution but is more soluble in many organic liquids (purple bottom layer).

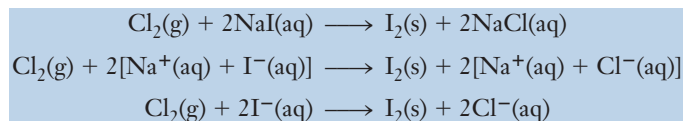
Charles D. Winters

**Plan**

The activity of the halogens decreases from top to bottom in the periodic table. We see (a) that Br is above I and (c) that Cl is above Br in the periodic table; therefore, neither combination (a) nor combination (c) could result in reaction. Cl is above I in the periodic table, and so combination (b) results in a displacement reaction.

**Solution**

The more active halogen, Cl<sub>2</sub>, displaces the less active halogen, I<sub>2</sub>, from its compounds.



You should now work Exercise 86.

**4-11 METATHESIS REACTIONS**

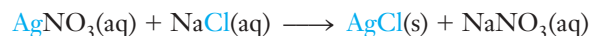
In many reactions between two compounds in aqueous solution, the positive and negative ions appear to “change partners” to form two new compounds, with no change in oxidation numbers. Such reactions are called **metathesis reactions**.

Pronounced “meh-tath-uh-sis.” Metathesis reactions are also sometimes referred to as **double displacement reactions**.

We can represent such reactions by the following general equation, where A and B represent positive ions (cations) and X and Y represent negative ions (anions):



For example, when we mix silver nitrate and sodium chloride solutions, solid silver chloride is formed and sodium nitrate remains dissolved in water:



Metathesis reactions result in the removal of ions from solution; this removal of ions can be thought of as the *driving force* for the reaction—the reason it occurs. The removal of ions can occur in three ways, which can be used to classify two types of metathesis reactions:

1. Formation of predominantly nonionized molecules (weak or nonelectrolytes) in solution; the most common such nonelectrolyte product is water
2. Formation of an insoluble solid, called a *precipitate* (which separates from the solution)

**1 Acid–Base (Neutralization) Reactions: Formation of a Nonelectrolyte**

Acid–base reactions are among the most important kinds of chemical reactions. Many acid–base reactions occur in nature in both plants and animals. Many acids and bases are essential compounds in an industrialized society (see Table 4-15). For example, approximately 300 pounds of sulfuric acid, H<sub>2</sub>SO<sub>4</sub>, and approximately 100 pounds of ammonia, NH<sub>3</sub>, is required to support the lifestyle of an average American for one year.

The reaction of an acid with a metal hydroxide base produces a salt and water. Such reactions are called **neutralization reactions** because the typical properties of acids and bases are neutralized.

The manufacture of fertilizers consumes more H<sub>2</sub>SO<sub>4</sub> and more NH<sub>3</sub> than any other single use.

**TABLE 4-15** 2001 Production of Inorganic Acids, Bases, and Salts in the United States

Formula	Name	Billions of Pounds	Major Uses
H <sub>2</sub> SO <sub>4</sub>	sulfuric acid	80.11	Manufacture of fertilizers and other chemicals
CaO, Ca(OH) <sub>2</sub>	lime (calcium oxide and calcium hydroxide)	41.23	Manufacture of other chemicals, steelmaking, water treatment
NH <sub>3</sub>	ammonia	26.09	Fertilizer; manufacture of fertilizers and other chemicals
H <sub>3</sub> PO <sub>4</sub>	phosphoric acid	23.21	Manufacture of fertilizers
Na <sub>2</sub> CO <sub>3</sub>	sodium carbonate (soda ash)	22.71	Manufacture of glass, other chemicals, detergents, pulp, and paper
NaOH	sodium hydroxide	22.37	Manufacture of other chemicals, pulp and paper, soap and detergents, aluminum, textiles
HNO <sub>3</sub>	nitric acid	15.65	Manufacture of fertilizers, explosives, plastics, and lacquers
NH <sub>4</sub> NO <sub>3</sub>	ammonium nitrate	14.20	Fertilizer and explosives
HCl	hydrochloric acid	8.84	Manufacture of other chemicals and rubber; metal cleaning
(NH <sub>4</sub> ) <sub>2</sub> SO <sub>4</sub>	ammonium sulfate	5.12	Fertilizer
TiO <sub>2</sub>	titanium dioxide	2.93	Paints and coatings
KOH, K <sub>2</sub> CO <sub>3</sub>	potash	2.65	Manufacture of fertilizers
Al <sub>2</sub> (SO <sub>4</sub> ) <sub>2</sub>	aluminum sulfate	2.33	Water treatment, dyeing textiles
NaClO <sub>3</sub>	sodium chlorate	1.81	Manufacture of other chemicals, explosives, plastics
Na <sub>2</sub> SO <sub>4</sub>	sodium sulfate	1.13	Manufacture of paper, glass, and detergents

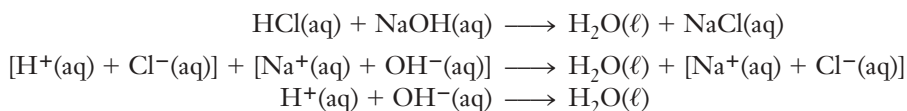
When a base such as ammonia or an amine reacts with an acid, a salt, but no water, is formed. This is still called an acid–base, or neutralization, reaction.

In nearly all neutralization reactions, the driving force is the combination of H<sup>+</sup>(aq) from an acid and OH<sup>-</sup>(aq) from a base (or a base plus water) to form water molecules.

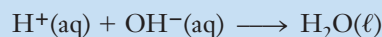
Consider the reaction of hydrochloric acid, HCl(aq), with aqueous sodium hydroxide, NaOH. Table 4-5 tells us that HCl is a strong acid, and Table 4-7 tells us that NaOH is a strong base. The salt sodium chloride, NaCl, is formed in this reaction. It contains the cation of its parent base, Na<sup>+</sup>, and the anion of its parent acid, Cl<sup>-</sup>. Solubility guidelines 2 and 4 tell us that NaCl is a soluble salt.



CD-ROM Screen 5.10, Acid–Base Reactions



The net ionic equation for *all* reactions of strong acids with strong bases that form soluble salts and water is



 **Problem-Solving Tip: Salt Formation**

The salt that is formed in a neutralization reaction is composed of the cation of the base and the anion of the acid. The salt may be soluble or insoluble. If our goal were to obtain the salt from the reaction of aqueous HCl with aqueous NaOH, we could evaporate the water and obtain solid NaCl.

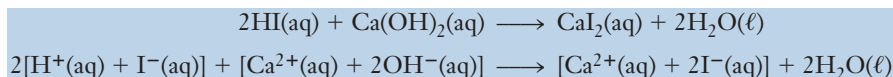
### EXAMPLE 4-10 Neutralization Reactions

Predict the products of the reaction between HI(aq) and Ca(OH)<sub>2</sub>(aq). Write balanced formula unit, total ionic, and net ionic equations.

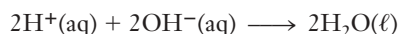
**Plan**

This is an acid–base neutralization reaction; the products are H<sub>2</sub>O and the salt that contains the cation of the base, Ca<sup>2+</sup>, and the anion of the acid, I<sup>−</sup>; CaI<sub>2</sub> is a soluble salt (solubility guideline 4). HI is a strong acid (see Table 4-5), Ca(OH)<sub>2</sub> is a strong base (see Table 4-7), and CaI<sub>2</sub> is a soluble ionic salt, so all are written in ionic form.

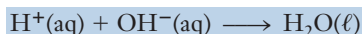
**Solution**



We cancel the spectator ions.



Dividing by 2 gives the net ionic equation:



*You should now work Exercise 91.*

Recall that in balanced equations we show the smallest whole-number coefficients possible.

Reactions of *weak* acids with strong bases also produce salts and water, but there is a significant difference in the balanced ionic equations because weak acids are only *slightly* ionized.

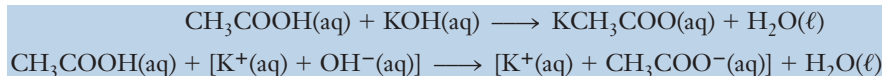
### EXAMPLE 4-11 Neutralization Reactions

Write balanced formula unit, total ionic, and net ionic equations for the reaction of acetic acid with potassium hydroxide.

**Plan**

Neutralization reactions involving metal hydroxide bases produce a salt and water. CH<sub>3</sub>COOH is a weak acid (see Table 4-6), so it is written as formula units. KOH is a strong base (see Table 4-7) and KCH<sub>3</sub>COO is a soluble salt (solubility guidelines 2 and 3), so both are written in ionic form.

**Solution**



The spectator ion is  $K^+$ , the cation of the strong base, KOH.

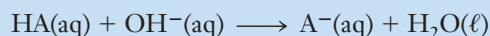


Thus, we see that *this* net ionic equation includes *molecules* of the weak acid and *anions* of the weak acid.

*You should now work Exercise 92.*

A *monoprotic acid* contains one acidic H per formula unit.

The reactions of *weak monoprotic acids* with *strong bases* that form *soluble salts* can be represented in general terms as



where HA represents the weak acid and  $A^-$  represents its anion.

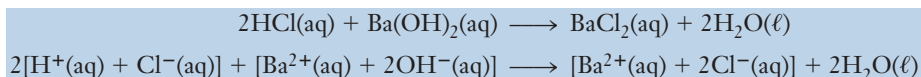
### EXAMPLE 4-12 Salt Formation

Write balanced formula unit, total ionic, and net ionic equations for an acid–base reaction that will produce the salt, barium chloride.

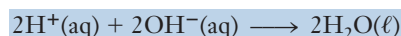
#### Plan

Neutralization reactions produce a salt. The salt contains the cation from the base and the anion from the acid. The base must therefore contain  $Ba^{2+}$ , that is,  $Ba(OH)_2$ , and the acid must contain  $Cl^-$ , that is, HCl. We write equations that represent the reaction between the strong base,  $Ba(OH)_2$ , and the strong acid, HCl.

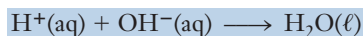
#### Solution



We cancel the spectator ions.



Dividing by 2 gives the net ionic equation:



*You should now work Exercise 100.*

The net ionic equation shows the driving force for this reaction. The formula unit equation shows the salt formed or that could be isolated if the water were evaporated.

## 2 Precipitation Reactions

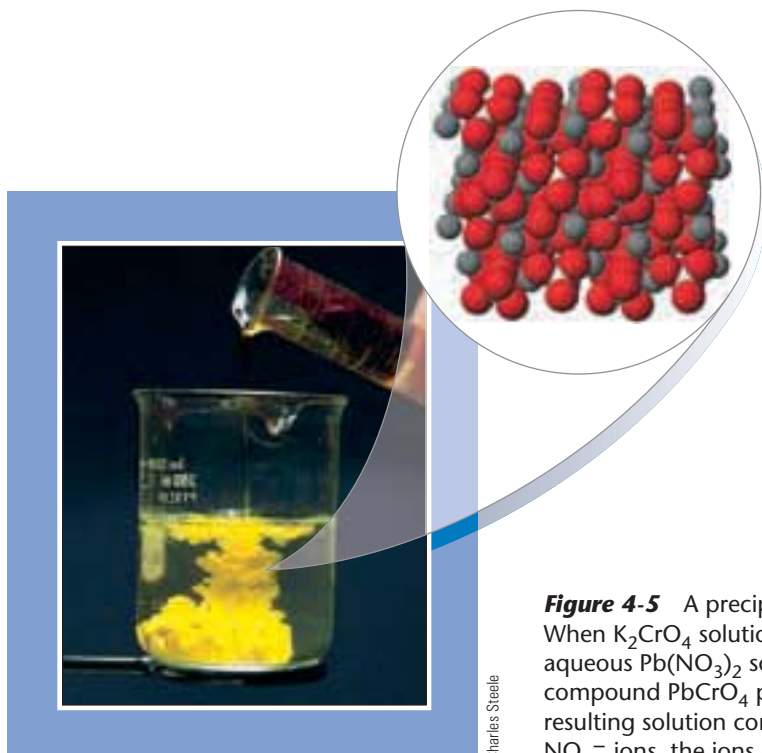
In **precipitation reactions** an insoluble solid, a **precipitate**, forms and then settles out of solution. The driving force for these reactions is the strong attraction between cations and anions. This results in the removal of ions from solution by the formation of a precipitate. Our teeth and bones were formed by very slow precipitation reactions in which mostly calcium phosphate  $Ca_3(PO_4)_2$  was deposited in the correct geometric arrangements.

An example of a precipitation reaction is the formation of bright yellow insoluble lead(II) chromate when we mix solutions of the soluble ionic compounds lead(II) nitrate and potassium chromate (Figure 4-5). The other product of the reaction is  $KNO_3$ , a soluble ionic salt.



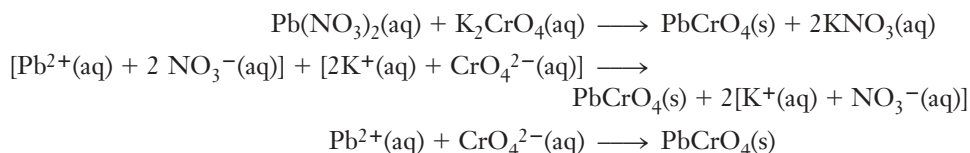
To understand the discussion of precipitation reactions, you *must* know the solubility guidelines, Table 4-8.





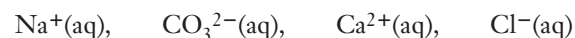
**Figure 4-5** A precipitation reaction. When  $\text{K}_2\text{CrO}_4$  solution is added to aqueous  $\text{Pb}(\text{NO}_3)_2$  solution, the yellow compound  $\text{PbCrO}_4$  precipitates. The resulting solution contains  $\text{K}^+$  and  $\text{NO}_3^-$  ions, the ions of  $\text{KNO}_3$ .

The balanced formula unit, total ionic, and net ionic equations for this reaction follow:

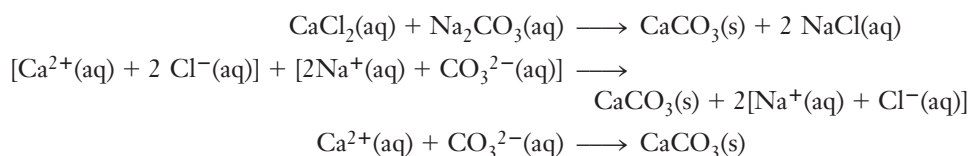


Another important precipitation reaction involves the formation of insoluble carbonates (solubility guideline 7). Limestone deposits are mostly calcium carbonate,  $\text{CaCO}_3$ , although many also contain significant amounts of magnesium carbonate,  $\text{MgCO}_3$ .

Suppose we mix together aqueous solutions of sodium carbonate,  $\text{Na}_2\text{CO}_3$ , and calcium chloride,  $\text{CaCl}_2$ . We recognize that *both*  $\text{Na}_2\text{CO}_3$  and  $\text{CaCl}_2$  (solubility guidelines 2, 4a, and 7) are soluble ionic compounds. At the instant of mixing, the resulting solution contains four ions:



One pair of ions,  $\text{Na}^+$  and  $\text{Cl}^-$ , *cannot* form an insoluble compound (solubility guidelines 2 and 4). We look for a pair of ions that could form an insoluble compound.  $\text{Ca}^{2+}$  ions and  $\text{CO}_3^{2-}$  ions are such a combination; they form insoluble  $\text{CaCO}_3$  (solubility guideline 7). The equations for the reaction follow:



Seashells, which are formed in very slow precipitation reactions, are mostly calcium carbonate ( $\text{CaCO}_3$ ), a white compound. Traces of transition metal ions give them color.



**TOPIC:** Precipitation Reactions  
**GO TO:** [www.scilinks.org](http://www.scilinks.org)  
**scILINKS CODE:** WCH0480

### EXAMPLE 4-13 Solubility Guidelines and Precipitation Reactions

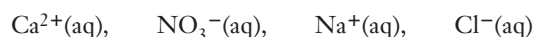
Will a precipitate form when aqueous solutions of  $\text{Ca}(\text{NO}_3)_2$  and  $\text{NaCl}$  are mixed in reasonable concentrations? Write balanced formula unit, total ionic, and net ionic equations for any reaction.

#### Plan

We recognize that both  $\text{Ca}(\text{NO}_3)_2$  (solubility guideline 3) and  $\text{NaCl}$  (solubility guidelines 2 and 4) are soluble compounds. We use the solubility guidelines to determine whether any of the possible products are insoluble.

#### Solution

At the instant of mixing, the resulting solution contains four ions:



New combinations of ions *could* be  $\text{CaCl}_2$  and  $\text{NaNO}_3$ . But solubility guideline 4 tells us that  $\text{CaCl}_2$  is a soluble compound, and solubility guidelines 2 and 3 tell us that  $\text{NaNO}_3$  is a soluble compound. Therefore, **no precipitate forms in this solution.**

*You should now work Exercise 104.*

### EXAMPLE 4-14 Solubility Guidelines and Precipitation Reactions

Will a precipitate form when aqueous solutions of  $\text{CaCl}_2$  and  $\text{K}_3\text{PO}_4$  are mixed in reasonable concentrations? Write balanced formula unit, total ionic, and net ionic equations for any reaction.

#### Plan

We recognize that both  $\text{CaCl}_2$  (solubility guideline 4) and  $\text{K}_3\text{PO}_4$  (solubility guideline 2) are soluble compounds. We use the solubility guidelines to determine whether any of the possible products are insoluble.

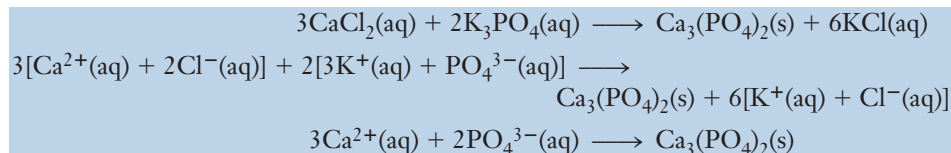
#### Solution

At the instant of mixing, the resulting solution contains four ions:



New combinations of ions *could* be  $\text{KCl}$  and  $\text{Ca}_3(\text{PO}_4)_2$ . Solubility guidelines 2 and 4 tell us that  $\text{KCl}$  is a soluble compound; solubility guideline 7 tells us that  $\text{Ca}_3(\text{PO}_4)_2$  is an insoluble compound, so **a precipitate of  $\text{Ca}_3(\text{PO}_4)_2$  forms in this solution.**

The equations for the formation of calcium phosphate follow:



*You should now work Exercise 110.*

**4-12 SUMMARY OF REACTION TYPES**

Table 4-16 summarizes the reaction types we have presented. Remember that a reaction might be classified in more than one category.



CD-ROM Screen 5.6, Types of Reactions in Aqueous Solution

**EXAMPLE 4-15 Classifying Reactions**

Classify each of the following reactions.

- (a)  $\text{Zn(s)} + \text{AgNO}_3\text{(aq)} \longrightarrow \text{Zn(NO}_3)_2\text{(aq)} + 2\text{Ag(s)}$   
 (b)  $\text{Ca(OH)}_2\text{(s)} \xrightarrow{\text{heat}} \text{CaO(s)} + \text{H}_2\text{O(g)}$   
 (c)  $2\text{HI(g)} \xrightarrow{\text{heat}} \text{H}_2\text{(g)} + \text{I}_2\text{(g)}$   
 (d)  $\text{Cu(NO}_3)_2\text{(aq)} + \text{Na}_2\text{S(aq)} \longrightarrow \text{CuS(s)} + 2\text{NaNO}_3\text{(aq)}$   
 (e)  $\text{SO}_2\text{(g)} + \text{H}_2\text{O(l)} \longrightarrow \text{H}_2\text{SO}_3\text{(aq)}$   
 (f)  $\text{H}_2\text{SO}_3\text{(aq)} + 2\text{KOH(aq)} \longrightarrow \text{K}_2\text{SO}_3\text{(aq)} + 2\text{H}_2\text{O(l)}$

**Plan**

We identify each reaction type by its characteristics, using Table 4-16 and the appropriate sections as a guide.

**Solution**

- (a) One element, Zn, displaces another, Ag, from a compound; this is a **displacement reaction**.  
 (b) A single compound breaks apart into two compounds; this is a **decomposition reaction**.  
 (c) A single compound breaks apart into two elements; this is another **decomposition reaction**. However, now there are changes in oxidation numbers; H changes from +1 in HI to 0 in H<sub>2</sub>, and I changes from -1 in HI to 0 in I<sub>2</sub>. So this is also an **oxidation-reduction (redox) reaction**.  
 (d) The positive and negative ions in the two reactant compounds change partners; this is a **metathesis reaction**. An insoluble product, CuS(s), is formed, so the reaction is a **precipitation reaction**.  
 (e) Two compounds combine to form a single product; this is a **combination reaction**.  
 (f) The positive and negative ions change partners; this is a **metathesis reaction**. An acid and a base react to form a salt and water; this is an **acid-base (neutralization) reaction**.

*You should now work Exercises 113 through 121.*

**TABLE 4-16** Summary and Examples of Reaction Types

Section Reaction Type, Examples	Characteristics
4-7 Oxidation–Reduction (redox)	Oxidation numbers (Section 4-4) of some elements change; at least one element must increase and at least one must decrease in oxidation number
4-8 Combination 1. element + element → compound $2\text{Al(s)} + 3\text{Cl}_2\text{(g)} \rightarrow 2\text{AlCl}_3\text{(s)}^*$ $\text{P}_4\text{(s)} + 10\text{Cl}_2\text{(g)} \rightarrow 4\text{PCl}_5\text{(s)}^*$ 2. compound + element → compound $\text{SF}_4\text{(g)} + \text{F}_2\text{(g)} \rightarrow \text{SF}_6\text{(g)}^*$ $2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2\text{SO}_3\text{(l)}^*$ 3. compound + compound → compound $\text{CaO(s)} + \text{CO}_2\text{(g)} \rightarrow \text{CaCO}_3\text{(s)}$ $\text{Na}_2\text{O(s)} + \text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)}$	More than one reactant, single product
4-9 Decomposition 1. compound → element + element $2\text{HgO(s)} \rightarrow 2\text{Hg(g)} + \text{O}_2\text{(g)}^*$ $2\text{H}_2\text{O(l)} \rightarrow 2\text{H}_2\text{(g)} + \text{O}_2\text{(g)}^*$ 2. compound → compound + element $2\text{NaNO}_3\text{(s)} \rightarrow 2\text{NaNO}_2\text{(s)} + \text{O}_2\text{(g)}^*$ $2\text{H}_2\text{O}_2\text{(l)} \rightarrow 2\text{H}_2\text{O(l)} + \text{O}_2\text{(g)}^*$ 3. compound → compound + compound $\text{CaCO}_3\text{(s)} \rightarrow \text{CaO(s)} + \text{CO}_2\text{(g)}$ $\text{Mg(OH)}_2\text{(s)} \rightarrow \text{MgO(s)} + \text{H}_2\text{O(l)}$	Single reactant, more than one product
4-10 Displacement $\text{Zn(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{Cu(s)} + \text{ZnSO}_4\text{(aq)}^*$ $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{H}_2\text{(g)} + \text{ZnSO}_4\text{(aq)}^*$ $\text{Cl}_2\text{(g)} + 2\text{NaI(aq)} \rightarrow \text{I}_2\text{(s)} + 2\text{NaCl(aq)}^*$	One element displaces another from a compound: Element + compound → element + compound Activity series (Table 4-14) summarizes metals and hydrogen; halogen activities (Group VIIA) decrease going down the group
4-11 Metathesis	Positive and negative ions in two compounds appear to “change partners” to form two new compounds; no change in oxidation numbers
1. acid–base neutralization $\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}$ $\text{CH}_3\text{COOH(aq)} + \text{KOH(aq)} \rightarrow \text{KCH}_3\text{COO(aq)} + \text{H}_2\text{O(l)}$ $\text{HCl(aq)} + \text{NH}_3\text{(aq)} \rightarrow \text{NH}_4\text{Cl(aq)}$ $2\text{H}_3\text{PO}_4\text{(aq)} + 3\text{Ca(OH)}_2\text{(aq)} \rightarrow \text{Ca}_3\text{(PO}_4)_2\text{(s)} + 6\text{H}_2\text{O(l)}^\dagger$	Product is a salt; water is often formed
2. precipitation $\text{CaCl}_2\text{(aq)} + \text{Na}_2\text{CO}_3\text{(aq)} \rightarrow \text{CaCO}_3\text{(s)} + 2\text{NaCl(aq)}$ $\text{Pb(NO}_3)_2\text{(aq)} + \text{K}_2\text{CrO}_4\text{(aq)} \rightarrow \text{PbCrO}_4\text{(s)} + 2\text{KNO}_3\text{(aq)}$ $2\text{H}_3\text{PO}_4\text{(aq)} + 3\text{Ca(OH)}_2\text{(aq)} \rightarrow \text{Ca}_3\text{(PO}_4)_2\text{(s)} + 6\text{H}_2\text{O(l)}^\dagger$	Products include an insoluble substance, which precipitates from solution as a solid; solubility guidelines assist in predicting, recognizing

\*These examples are also oxidation–reduction (redox) reactions.

†This reaction is both an acid–base neutralization reaction and a precipitation reaction.

## Key Terms

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- Acid** A substance that produces  $\text{H}^+(\text{aq})$  ions in aqueous solution. Strong acids ionize completely or almost completely in dilute aqueous solution. Weak acids ionize only slightly.
- Acid–base reaction** See *Neutralization reaction*.
- Active metal** A metal that readily loses electrons to form cations.
- Activity series** A listing of metals (and hydrogen) in order of decreasing activity.
- Alkali metals** Elements of Group IA in the periodic table, except hydrogen.
- Alkaline earth metals** Group IIA elements in the periodic table.
- Atomic number** The number of protons in the nucleus of an atom of an element.
- Base** A substance that produces  $\text{OH}^-(\text{aq})$  ions in aqueous solution. Strong bases are soluble in water and are completely *dissociated*. Weak bases ionize only slightly.
- Binary acid** A binary compound in which H is bonded to a nonmetal in Group VIIA or a nonmetal other than oxygen in Group VIA.
- Binary compound** A compound consisting of two elements; may be ionic or molecular.
- Chemical periodicity** The variation in properties of elements with their positions in the periodic table.
- Combination reaction** Reaction in which two substances (elements or compounds) combine to form one compound.
- Decomposition reaction** Reaction in which a compound decomposes to form two or more products (elements, compounds, or some combination of these).
- Displacement reaction** A reaction in which one element displaces another from a compound.
- Disproportionation reaction** A redox reaction in which the oxidizing agent and the reducing agent are the same element.
- Dissociation** In aqueous solution, the process in which a solid *ionic compound* separates into its ions.
- Electrolyte** A substance whose aqueous solutions conduct electricity.
- Formula unit equation** An equation for a chemical reaction in which all formulas are written as complete formulas.
- Group (family)** The elements in a vertical column of the periodic table.
- Halogens** Group VIIA elements in the periodic table.
- Ionization** In aqueous solution, the process in which a *molecular compound* separates to form ions.
- Metal** An element below and to the left of the stepwise division (metalloids) of the periodic table; about 80% of the known elements are metals.
- Metalloids** Elements with properties intermediate between metals and nonmetals: B, Si, Ge, As, Sb, Te, Po, and At.
- Metathesis reaction** A reaction in which the positive and negative ions in two compounds “change partners,” with no change in oxidation numbers, to form two new compounds.
- Net ionic equation** An equation that results from canceling spectator ions from a total ionic equation.
- Neutralization reaction** The reaction of an acid with a base to form a salt. Often, the reaction of hydrogen ions with hydroxide ions to form water molecules.
- Noble (rare) gases** Elements of Group VIIIA in the periodic table.
- Nonelectrolyte** A substance whose aqueous solutions do not conduct electricity.
- Nonmetals** Elements above and to the right of the metalloids in the periodic table.
- Oxidation** An increase in oxidation number; corresponds to a loss of electrons.
- Oxidation numbers** Arbitrary numbers that can be used as mechanical aids in writing formulas and balancing equations; for single-atom ions they correspond to the charge on the ion; less metallic atoms are assigned negative oxidation numbers in compounds and polyatomic ions.
- Oxidation–reduction reaction** A reaction in which oxidation and reduction occur; also called a redox reaction.
- Oxidation states** See *Oxidation numbers*.
- Oxidizing agent** The substance that oxidizes another substance and is reduced.
- Oxoacid** See *Ternary acid*.
- Period** The elements in a horizontal row of the periodic table.
- Periodic law** The properties of the elements are periodic functions of their atomic numbers.
- Periodic table** An arrangement of elements in order of increasing atomic number that also emphasizes periodicity.
- Periodicity** Regular periodic variations of properties of elements with atomic number (and position in the periodic table).
- Precipitate** An insoluble solid that forms and separates from a solution.
- Precipitation reaction** A reaction in which a solid (precipitate) forms.
- Pseudobinary ionic compound** A compound that contains more than two elements but is named like a binary compound.
- Redox reaction** See *Oxidation–reduction reaction*.
- Reducing agent** The substance that reduces another substance and is oxidized.
- Reduction** A decrease in oxidation number; corresponds to a gain of electrons.
- Reversible reaction** A reaction that occurs in both directions; described with double arrows ( $\rightleftharpoons$ ).
- Salt** A compound that contains a cation other than  $\text{H}^+$  and an anion other than  $\text{OH}^-$  or  $\text{O}^{2-}$ .
- Semiconductor** A substance that does not conduct electricity at low temperatures but does so at higher temperatures.
- Spectator ions** Ions that appear in solution on both sides of the total ionic equation; they undergo no change in the chemical reaction.
- Strong acid** An acid that ionizes (separates into ions) completely, or very nearly completely, in dilute aqueous solution.

**Strong electrolyte** A substance that conducts electricity well in dilute aqueous solution.

**Strong base** Metal hydroxide that is soluble in water and dissociates completely in dilute aqueous solution.

**Ternary acid** A ternary compound containing H, O, and another element, usually a nonmetal.

**Ternary compound** A compound consisting of three elements; may be ionic or molecular.

**Ternary salt** A salt resulting from replacing the hydrogen in a ternary acid with another ion.

**Total ionic equation** An equation for a chemical reaction written to show the predominant form of all species in aqueous solution or in contact with water.

**Weak acid** An acid that ionizes only slightly in dilute aqueous solution.

**Weak base** A molecular substance that ionizes only slightly in water to produce an alkaline (base) solution.

**Weak electrolyte** A substance that conducts electricity poorly in dilute aqueous solution.

## Exercises

### The Periodic Table

1. State the periodic law. What does it mean?
2. What was Mendeleev's contribution to the construction of the modern periodic table?
3. Consult a handbook of chemistry, and look up melting points of the elements of Periods 2 and 3. Show that melting point is a property that varies periodically for these elements.
- \*4. Mendeleev's periodic table was based on increasing atomic weight. Argon has a higher atomic weight than potassium, yet in the modern table argon appears before potassium. Explain how this can be.
5. Estimate the density of antimony from the following densities ( $\text{g}/\text{cm}^3$ ): As, 5.72; Bi, 9.8; Sn, 7.30; Te, 6.24. Show how you arrived at your answer. Using a reference other than your textbook, look up the density of antimony. How does your predicted value compare with the reported value?



Charles D. Winters

Product made from antimony

6. Estimate the density of selenium from the following densities ( $\text{g}/\text{cm}^3$ ): S, 2.07; Te, 6.24; As, 5.72; Br, 3.12. Show how you arrived at your answer. Using a reference other than your textbook, look up the density of selenium. How does your predicted value compare with the reported value?
7. Estimate the specific heat of antimony from the following specific heats ( $\text{J}/\text{g}^\circ\text{C}$ ): As, 0.34; Bi, 0.14; Sn, 0.23; Te, 0.20. Show how you arrived at your answer.

8. Given the following melting points in  $^\circ\text{C}$ , estimate the value for  $\text{CBr}_4$ :  $\text{CF}_4$ ,  $-184$ ;  $\text{CCl}_4$ ,  $-23$ ;  $\text{CI}_4$ ,  $171$  (decomposes). Using a reference other than your textbook, look up the melting point of  $\text{CBr}_4$ . How does your predicted value compare with the reported value?
9. Calcium and magnesium form the following compounds:  $\text{CaCl}_2$ ,  $\text{MgCl}_2$ ,  $\text{CaO}$ ,  $\text{MgO}$ ,  $\text{Ca}_3\text{N}_2$ , and  $\text{Mg}_3\text{N}_2$ . Predict the formula for a compound of (a) magnesium and sulfur, (b) barium and iodine.
10. The formulas of some hydrides of second-period representative elements are as follows:  $\text{BeH}_2$ ,  $\text{BH}_3$ ,  $\text{CH}_4$ ,  $\text{NH}_3$ ,  $\text{H}_2\text{O}$ ,  $\text{HF}$ . A famous test in criminology laboratories for the presence of arsenic (As) involves the formation of arsine, the hydride of arsenic. Predict the formula of arsine.



Charles Steele

Arsine burns to form a dark spot.

11. Clearly distinguish between the following terms and provide specific examples of each: groups (families) of elements, and periods of elements.
12. Write names and symbols for (a) the alkaline earth metals, (b) the Group IVA elements, (c) the Group VIB elements.
13. Write names and symbols for (a) the alkali metals, (b) the noble gases, (c) the Group IIIA elements.
14. Clearly and concisely define the following terms and



provide examples of each: (a) metals, (b) nonmetals, (c) halogens.

### Aqueous Solutions

- Define and distinguish among (a) strong electrolytes, (b) weak electrolytes, and (c) nonelectrolytes.
- Three common classes of compounds are electrolytes. Name them and give an example of each.
- Define (a) acids, (b) bases, (c) salts, and (d) molecular compounds.
- How can a salt be related to a particular acid and a particular base?
- List the names and formulas of seven strong acids.
- Write equations for the ionization of the following acids: (a) hydrochloric acid, (b) nitric acid, (c) perchloric acid.
- List the names and formulas of five weak acids.
- List the names and formulas of the common soluble strong bases.
- Write equations for the ionization of the following acids. Which ones ionize only slightly? (a) HF, (b)  $\text{H}_2\text{SO}_4$ , (c)  $\text{CH}_3\text{COOH}$ , (d)  $\text{HNO}_2$ .
- The most common weak base is present in a common household chemical. Write the equation for the ionization of this weak base.
- Summarize the electrical properties of strong electrolytes, weak electrolytes, and nonelectrolytes.
- What is the difference between ionization and dissociation in aqueous solution?
- List the names and formulas of five insoluble bases.
- Which of the following are strong electrolytes? Weak electrolytes? Nonelectrolytes? (a)  $\text{Na}_2\text{S}$ , (b)  $\text{Ba}(\text{OH})_2$ , (c)  $\text{CH}_3\text{OH}$ , (d)  $\text{HCN}$ , (e)  $\text{Al}(\text{NO}_3)_3$ .
- Classify the following as strong electrolytes, weak electrolytes, or nonelectrolytes: (a)  $\text{NaClO}_4$ , (b)  $\text{HClO}_2$ , (c)  $\text{CH}_3\text{CH}_2\text{OH}$ , (d)  $\text{CH}_3\text{COOH}$ , (e)  $\text{HNO}_3$ .
- Write the formulas of two soluble and two insoluble chlorides, sulfates, and hydroxides.
- Describe an experiment for classifying each of these compounds as a strong electrolyte, a weak electrolyte, or a nonelectrolyte:  $\text{Na}_2\text{CO}_3$ ,  $\text{HCN}$ ,  $\text{CH}_3\text{COOH}$ ,  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{NH}_3$ . Predict and explain the expected results.
- (a) Which of these are acids?  $\text{HI}$ ,  $\text{NH}_3$ ,  $\text{H}_2\text{SeO}_4$ ,  $\text{BF}_3$ ,  $\text{Fe}(\text{OH})_3$ ,  $\text{H}_2\text{S}$ ,  $\text{C}_6\text{H}_6$ ,  $\text{CsOH}$ ,  $\text{H}_3\text{PO}_3$ ,  $\text{HCN}$ . (b) Which of these are bases?  $\text{NaOH}$ ,  $\text{H}_2\text{Se}$ ,  $\text{BCl}_3$ ,  $\text{NH}_3$ .
- \*Classify each substance as either an electrolyte or a nonelectrolyte:  $\text{NH}_4\text{Cl}$ ,  $\text{HI}$ ,  $\text{C}_6\text{H}_6$ ,  $\text{Zn}(\text{CH}_3\text{COO})_2$ ,  $\text{Cu}(\text{NO}_3)_2$ ,  $\text{CH}_3\text{CH}_2\text{OH}$ ,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$  (sugar),  $\text{LiOH}$ ,  $\text{KHCO}_3$ ,  $\text{CCl}_4$ ,  $\text{La}_2(\text{SO}_4)_3$ ,  $\text{I}_2$ .
- \*Classify each substance as either a strong or weak electrolyte, and then list (a) the strong acids, (b) the strong bases, (c) the weak acids, and (d) the weak bases.  $\text{NaBr}$ ,  $\text{MgSO}_4$ ,  $\text{HCl}$ ,  $\text{H}_2\text{C}_2\text{O}_4$ ,  $\text{Ba}(\text{NO}_3)_2$ ,  $\text{H}_3\text{PO}_4$ ,  $\text{CsOH}$ ,  $\text{HNO}_3$ ,  $\text{HI}$ ,  $\text{Ba}(\text{OH})_2$ ,  $\text{LiOH}$ ,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{KOH}$ ,  $\text{Mg}(\text{CH}_3\text{COO})_2$ ,  $\text{HCN}$ ,  $\text{HClO}_4$ .

- Classify each substance as soluble, moderately soluble, or insoluble.  $\text{Ag}_2\text{SO}_4$ ,  $(\text{NH}_4)_2\text{CO}_3$ ,  $\text{AgCl}$ ,  $\text{HgBr}_2$
- Classify each substance as soluble, moderately soluble, or insoluble.  $\text{Ca}(\text{CH}_3\text{COO})_2$ ,  $\text{NH}_4\text{Cl}$ ,  $\text{AgNO}_3$ ,  $\text{PbCl}_2$
- What are reversible reactions? Give some examples.
- Many household “chemicals” are acidic or basic. List a few of each kind.
- Some chemical reactions reach an equilibrium rather than going to completion. What is “equal” in such an equilibrium?
- Vinegar is 5% acetic acid, an organic acid, by mass. Many organic acids occur in living systems. What conclusion can be drawn from this information as to the strengths of organic acids?

### Oxidation Numbers

- Assign oxidation numbers to the element specified in each group of compounds. (a) N in  $\text{NO}$ ,  $\text{N}_2\text{O}_3$ ,  $\text{N}_2\text{O}_4$ ,  $\text{NH}_4\text{Cl}$ ,  $\text{N}_2\text{H}_4$ ,  $\text{NH}_2\text{OH}$ ,  $\text{HNO}_2$ ,  $\text{HNO}_3$  (b) C in  $\text{CO}$ ,  $\text{CO}_2$ ,  $\text{CH}_2\text{O}$ ,  $\text{CH}_4\text{O}$ ,  $\text{C}_2\text{H}_6\text{O}$ ,  $(\text{COOH})_2$ ,  $\text{Na}_2\text{CO}_3$ ,  $\text{C}_6\text{H}_6$  (c) S in  $\text{S}^{2-}$ ,  $\text{SO}_3^{2-}$ ,  $\text{SO}_4^{2-}$ ,  $\text{S}_2\text{O}_3^{2-}$ ,  $\text{S}_4\text{O}_6^{2-}$ ,  $\text{HS}$
- Assign oxidation numbers to the element specified in each group of compounds. (a) P in  $\text{PCl}_5$ ,  $\text{P}_4\text{O}_6$ ,  $\text{P}_4\text{O}_{10}$ ,  $\text{HPO}_3$ ,  $\text{H}_3\text{PO}_3$ ,  $\text{POCl}_3$ ,  $\text{H}_4\text{P}_2\text{O}_7$ ,  $\text{Mg}_3(\text{PO}_4)_2$  (b) Br in  $\text{Br}^-$ ,  $\text{BrO}^-$ ,  $\text{BrO}_2^-$ ,  $\text{BrO}_3^-$ ,  $\text{BrO}_4^-$  (c) Mn in  $\text{MnO}$ ,  $\text{MnO}_2$ ,  $\text{Mn}(\text{OH})_2$ ,  $\text{K}_2\text{MnO}_4$ ,  $\text{KMnO}_4$ ,  $\text{Mn}_2\text{O}_7$  (d) O in  $\text{OF}_2$ ,  $\text{Na}_2\text{O}$ ,  $\text{Na}_2\text{O}_2$ ,  $\text{KO}_2$ .
- Assign oxidation numbers to the element specified in each group of ions. (a) S in  $\text{S}_8$ ,  $\text{H}_2\text{S}$ ,  $\text{SO}_2$ ,  $\text{SO}_3$ ,  $\text{Na}_2\text{SO}_3$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{K}_2\text{SO}_4$  (b) Cr in  $\text{CrO}_2^-$ ,  $\text{Cr}(\text{OH})_4^-$ ,  $\text{CrO}_4^-$ ,  $\text{Cr}_2\text{O}_7^{2-}$  (c) B in  $\text{BO}_2^-$ ,  $\text{BO}_3^{3-}$ ,  $\text{B}_4\text{O}_7^{2-}$ .
- Assign oxidation numbers to the element specified in each group of ions. (a) N in  $\text{N}^{3-}$ ,  $\text{NO}_2^-$ ,  $\text{NO}_3^-$ ,  $\text{N}_3^-$ ,  $\text{NH}_4^-$  (b) Cl in  $\text{Cl}_2$ ,  $\text{HCl}$ ,  $\text{HClO}$ ,  $\text{HClO}_2$ ,  $\text{KClO}_3$ ,  $\text{Cl}_2\text{O}_7$ ,  $\text{Ca}(\text{ClO}_4)_2$ ,  $\text{PCl}_5^-$ .

### Naming Compounds

- Name the following common anions using the IUPAC system of nomenclature: (a)  $\text{NO}_3^-$ ; (b)  $\text{SO}_4^{2-}$ ; (c)  $\text{ClO}_3^-$ ; (d)  $\text{CH}_3\text{COO}^-$ ; (e)  $\text{PO}_4^{2-}$ .
- Name the following monatomic cations using the IUPAC system of nomenclature: (a)  $\text{Na}^+$ , (b)  $\text{Au}^{3+}$ , (c)  $\text{Ca}^{2+}$ , (d)  $\text{Zn}^{2+}$ , (e)  $\text{Ag}^+$ .
- Write the chemical symbol for each of the following: (a) sodium ion, (b) iron(II) ion, (c) silver ion, (d) mercury(II) ion, (e) bismuth(III) ion.
- Write the chemical formula for each of the following: (a) chloride ion, (b) hydrogen sulfide ion, (c) telluride ion, (d) oxide ion, (e) nitrite ion.
- Name the following ionic compounds: (a)  $\text{Li}_2\text{S}$ , (b)  $\text{SnO}_2$ , (c)  $\text{RbBr}$ , (d)  $\text{K}_2\text{O}$ , (e)  $\text{Ba}_3\text{N}_2$ .
- Name the following ionic compounds: (a)  $\text{CuI}_2$ , (b)  $\text{Hg}_2\text{Cl}_2$ , (c)  $\text{Li}_3\text{N}$ , (d)  $\text{MnCl}_2$ , (e)  $\text{CuCO}_3$ , (f)  $\text{FeO}$ .

51. Write the chemical formula for each of the following compounds: (a) lithium fluoride, (b) zinc oxide, (c) barium oxide, (d) magnesium bromide, (e) hydrogen cyanide, (f) copper(I) chloride.
52. Write the chemical formula for each of the following compounds: (a) copper(II) chlorite, (b) potassium nitrate, (c) barium phosphate, (d) copper(I) sulfate, (e) sodium sulfite.
53. What is the name of the acid with the formula  $\text{H}_2\text{CO}_3$ ? Write the formulas of the two anions derived from it and name these ions.
54. What is the name of the acid with the formula  $\text{H}_3\text{PO}_3$ ? What is the name of the  $\text{HPO}_3^{2-}$  ion?
55. Name the following binary molecular compounds: (a)  $\text{NO}_2$ , (b)  $\text{CO}_2$ , (c)  $\text{SF}_6$ , (d)  $\text{SiCl}_4$ , (e)  $\text{IF}$ .
56. Name the following binary molecular compounds: (a)  $\text{AsF}_3$ , (b)  $\text{Br}_2\text{O}$ , (c)  $\text{BrF}_5$ , (d)  $\text{CSe}_2$ , (e)  $\text{Cl}_2\text{O}_7$ .
57. Write the chemical formula for each of the following compounds: (a) iodine bromide, (b) silicon dioxide, (c) phosphorus trichloride, (d) tetrasulfur dinitride, (e) bromine trifluoride, (f) hydrogen telluride, (g) xenon tetrafluoride.
58. Write the chemical formula for each of the following compounds: (a) diboron trioxide, (b) dinitrogen pentasulfide, (c) phosphorus triiodide, (d) sulfur tetrachloride, (e) silicon sulfide, (f) hydrogen sulfide, (g) tetraphosphorus hexoxide.
59. Write formulas for the compounds that are expected to be formed by the following pairs of ions:

	A. $\text{Cl}^-$	B. $\text{OH}^-$	C. $\text{SO}_4^{2-}$	D. $\text{PO}_4^{3-}$	E. $\text{NO}_3^-$
1. $\text{NH}_4^+$		Omit – see note			
2. $\text{Na}^+$					
3. $\text{Mg}^{2+}$					
4. $\text{Ni}^{2+}$					
5. $\text{Fe}^{3+}$					
6. $\text{Ag}^+$					

NOTE: The compound  $\text{NH}_4\text{OH}$  does not exist. The solution commonly labeled “ $\text{NH}_4\text{OH}$ ” is aqueous ammonia,  $\text{NH}_3(\text{aq})$ .

60. Write the names for the compounds of Exercise 59.
61. Write balanced chemical equations for each of the following processes: (a) Calcium phosphate reacts with sulfuric acid to produce calcium sulfate and phosphoric acid. (b) Calcium phosphate reacts with water containing dissolved carbon dioxide to produce calcium hydrogen carbonate and calcium hydrogen phosphate.
62. Write balanced chemical equations for each of the following processes: (a) When heated, nitrogen and oxygen combine to form nitrogen oxide. (b) Heating a mixture of

lead(II) sulfide and lead(II) sulfate produces metallic lead and sulfur dioxide.

### Oxidation–Reduction Reactions

63. Define and provide examples of the following terms: (a) oxidation, (b) reduction, (c) oxidizing agent, (d) reducing agent.
64. Why must oxidation and reduction always occur simultaneously in chemical reactions?
65. Determine which of the following are oxidation–reduction reactions. For those that are, identify the oxidizing and reducing agents.
- (a)  $3\text{Zn}(\text{s}) + 2\text{CoCl}_3(\text{aq}) \rightarrow 3\text{ZnCl}_2(\text{aq}) + 2\text{Co}(\text{s})$   
 (b)  $\text{ICl}(\text{s}) + \text{H}_2\text{O}(\ell) \rightarrow \text{HCl}(\text{aq}) + \text{HIO}(\text{aq})$   
 (c)  $3\text{HCl}(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + \text{NOCl}(\text{g}) + 2\text{H}_2\text{O}(\ell)$   
 (d)  $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \xrightarrow{\text{heat}} 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$
66. Determine which of the following are oxidation–reduction reactions. For those that are, identify the oxidizing and reducing agents. (a)  $\text{HgCl}_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow \text{HgI}_2(\text{s}) + 2\text{KCl}(\text{aq})$  (b)  $4\text{NH}_3(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$  (c)  $\text{CaCO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Ca}(\text{NO}_3)_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$  (d)  $\text{PCl}_3(\ell) + 3\text{H}_2\text{O}(\ell) \rightarrow 3\text{HCl}(\text{aq}) + \text{H}_3\text{PO}_3(\text{aq})$
67. Write balanced formula unit equations for the following redox reactions: (a) aluminum reacts with sulfuric acid,  $\text{H}_2\text{SO}_4$ , to produce aluminum sulfate,  $\text{Al}_2(\text{SO}_4)_3$ , and hydrogen (b) nitrogen,  $\text{N}_2$ , reacts with hydrogen,  $\text{H}_2$ , to form ammonia,  $\text{NH}_3$  (c) zinc sulfide,  $\text{ZnS}$ , reacts with oxygen,  $\text{O}_2$ , to form zinc oxide,  $\text{ZnO}$ , and sulfur dioxide,  $\text{SO}_2$  (d) carbon reacts with nitric acid,  $\text{HNO}_3$ , to produce nitrogen dioxide,  $\text{NO}_2$ , carbon dioxide,  $\text{CO}_2$ , and water (e) sulfuric acid reacts with hydrogen iodide,  $\text{HI}$ , to produce sulfur dioxide,  $\text{SO}_2$ , iodine,  $\text{I}_2$ , and water.
68. Identify the oxidizing agents and reducing agents in the oxidation–reduction reactions given in Exercise 67.
69. Write total ionic and net ionic equations for the following redox reactions occurring in aqueous solution or in contact with water:
- (a)  $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$   
 (b)  $2\text{KMnO}_4 + 16\text{HCl} \rightarrow 2\text{MnCl}_2 + 2\text{KCl} + 5\text{Cl}_2 + 8\text{H}_2\text{O}$   
 (Note:  $\text{MnCl}_2$  is water-soluble.)  
 (c)  $4\text{Zn} + 10\text{HNO}_3 \rightarrow 4\text{Zn}(\text{NO}_3)_2 + \text{NH}_4\text{NO}_2 + 3\text{H}_2\text{O}$ .
70. Write total ionic and net ionic equations for the following unbalanced redox reactions occurring in aqueous solution or in contact with water:
- (a)  $\text{H}_2(\text{g}) + \text{Al}_2\text{O}_3(\text{s}) \rightarrow \text{Al}(\text{s}) + \text{H}_2\text{O}(\ell)$   
 (b)  $\text{Zn}(\text{s}) + \text{CoCl}_3(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{Co}(\text{s})$   
 (c)  $\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Fe}(\text{s}) + \text{CO}_2(\text{g})$ .

### Combination Reactions

71. Write balanced equations that show the combination reactions of the following Group IA metals combining with the

Group VIIA nonmetals. (a) Li and  $\text{Cl}_2$ ; (b) K and  $\text{F}_2$ ; (c) Na and  $\text{I}_2$

72. Write balanced equations that show the combination reactions of the following Group IIA metals and Group VIIA nonmetals. (a) Be and  $\text{F}_2$ , (b) Ca and  $\text{Br}_2$ , (c) Ba and  $\text{Cl}_2$

In Exercises 73 and 74, some combination reactions are described by words. Write the balanced chemical equation for each, and assign oxidation numbers to elements other than H and O.

73. (a) Antimony reacts with a limited amount of chlorine to form antimony(III) chloride. (b) Antimony(III) chloride reacts with excess chlorine to form antimony(V) chloride. (c) Carbon burns in a limited amount of oxygen to form carbon monoxide.
74. (a) Sulfur trioxide reacts with aluminum oxide to form aluminum sulfate. (b) Dichlorine heptoxide reacts with water to form perchloric acid. (c) When cement “sets,” the main reaction is the combination of calcium oxide with silicon dioxide to form calcium silicate,  $\text{CaSiO}_3$ .

### Decomposition Reactions

In Exercises 75 and 76, write balanced formula unit equations for the reactions described by words. Assign oxidation numbers to all elements.

75. (a) Hydrogen peroxide,  $\text{H}_2\text{O}_2$ , is used as an antiseptic. Blood causes it to decompose into water and oxygen; (b) When heated, ammonium nitrate can decompose explosively to form nitrogen oxide and steam.



A Model of Hydrogen Peroxide

76. (a) A “classroom volcano” is made by heating solid ammonium dichromate,  $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ , which decomposes into nitrogen, chromium(III) oxide, and steam. (b) At high temperatures, sodium nitrate (a fertilizer) forms sodium nitrite and oxygen.



Charles D. Winters

Ammonium Dichromate Volcano

### Displacement Reactions

77. Which of the following would displace hydrogen when a piece of the metal is dropped into dilute  $\text{H}_2\text{SO}_4$  solution? Write balanced total ionic and net ionic equations for the reactions: Zn, Cu, Sn, Al.
78. Which of the following metals would displace copper from an aqueous solution of copper(II) sulfate? Write balanced total ionic and net ionic equations for the reactions: Hg, Zn, Fe, Ni.
79. Arrange the metals listed in Exercise 77 in order of increasing activity.
80. Arrange the metals listed in Exercise 78 in order of increasing activity.
81. Which of the following metals would displace hydrogen from cold water? Write balanced net ionic equations for the reactions: Ag, Na, Ca, Cr.
82. Arrange the metals listed in Exercise 81 in order of increasing activity.
83. What is the order of decreasing activity of the halogens?
84. Name five elements that will react with steam but not with cold water.
85. (a) Name two common metals: one that *does not* displace hydrogen from water, and one that *does not* displace hydrogen from water or acid solutions. (b) Name two common metals: one that *does* displace hydrogen from water, and one that displaces hydrogen from acid solutions but not from water. Write net ionic equations for the reactions that occur.
86. Of the possible displacement reactions shown, which one(s) could occur?  
 (a)  $2\text{Cl}^-(\text{aq}) + \text{Br}_2(\ell) \rightarrow 2\text{Br}^-(\text{aq}) + \text{Cl}_2(\text{g})$   
 (b)  $2\text{Br}_2(\text{aq}) + \text{F}_2(\text{g}) \rightarrow 2\text{F}^-(\text{aq}) + \text{Br}_2(\ell)$   
 (c)  $2\text{I}^-(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{Cl}^-(\text{aq}) + \text{I}_2(\text{s})$   
 (d)  $2\text{Br}^-(\text{aq}) + \text{I}_2(\text{s}) \rightarrow 2\text{I}^-(\text{aq}) + \text{Br}_2(\ell)$
87. Use the activity series to predict whether or not the following reactions will occur:  
 (a)  $\text{Cu}(\text{s}) + \text{Mg}^+ \rightarrow \text{Mg}(\text{s}) + \text{Cu}^+$

- (b)  $\text{Ni(s)} + \text{Cu}^+ \rightarrow \text{Ni}^+ + \text{Cu(s)}$   
 (c)  $\text{Cu(s)} + 2\text{H}^+ \rightarrow \text{Cu}^+ + \text{H}_2(\text{g})$   
 (d)  $\text{Mg(s)} + \text{H}_2\text{O}(\text{g}) \rightarrow \text{MgO(s)} + \text{H}_2(\text{g})$

88. Repeat Exercise 87 for

- (a)  $\text{Sn(s)} + \text{Ca}^+ \rightarrow \text{Sn}^+ + \text{Ca(s)}$   
 (b)  $\text{Al}_2\text{O}_3(\text{s}) + 3\text{H}_2(\text{g}) \xrightarrow{\text{heat}} 2\text{Al(s)} + 3\text{H}_2\text{O(g)}$   
 (c)  $\text{Cu(s)} + 2\text{H}^+ \rightarrow \text{Cu}^+ + \text{H}_2(\text{g})$   
 (d)  $\text{Cu(s)} + \text{Pb}^+ \rightarrow \text{Cu}^+ + \text{Pb(s)}$

### Metathesis Reactions

Exercises 89 and 90 describe precipitation reactions *in aqueous solutions*. For each, write balanced (i) formula unit, (ii) total ionic, and (iii) net ionic equations. Refer to the solubility guidelines as necessary.

89. (a) Black-and-white photographic film contains some silver bromide, which can be formed by the reaction of sodium bromide with silver nitrate. (b) Barium sulfate is used when x-rays of the gastrointestinal tract are made. Barium sulfate can be prepared by reacting barium chloride with dilute sulfuric acid. (c) In water purification systems small solid particles are often “trapped” as aluminum hydroxide precipitates and fall to the bottom of the sedimentation pool. Aluminum sulfate reacts with calcium hydroxide (from lime) to form aluminum hydroxide and calcium sulfate.
90. (a) Our bones are mostly calcium phosphate. Calcium chloride reacts with potassium phosphate to form calcium phosphate and potassium chloride. (b) Mercury compounds are very poisonous. Mercury(II) nitrate reacts with sodium sulfide to form mercury(II) sulfide, which is very insoluble, and sodium nitrate. (c) Chromium(III) ions are very poisonous. They can be removed from solution by precipitating very insoluble chromium(III) hydroxide. Chromium(III) chloride reacts with calcium hydroxide to form chromium(III) hydroxide and calcium chloride.

In Exercises 91 through 94, write balanced (i) formula unit, (ii) total ionic, and (iii) net ionic equations for the reactions that occur between the acid and the base. Assume that all reactions occur in water or in contact with water.

91. (a) hydrochloric acid + calcium hydroxide (b) dilute sulfuric acid + strontium hydroxide (c) perchloric acid + aqueous ammonia
92. (a) acetic acid + potassium hydroxide (b) sulfurous acid + sodium hydroxide (c) hydrofluoric acid + sodium hydroxide
- \*93. (a) potassium hydroxide + hydrosulfuric acid (b) potassium hydroxide + hydrochloric acid (c) lead(II) hydroxide + hydrosulfuric acid
94. (a) lithium hydroxide + sulfuric acid (b) calcium hydroxide + phosphoric acid (c) copper(II) hydroxide + nitric acid
- In Exercises 95 through 98, write balanced (i) formula unit, (ii) total ionic, and (iii) net ionic equations for the reaction of an acid and a base that will produce the indicated salts.
95. (a) sodium chloride, (b) sodium phosphate, (c) barium nitrate

96. (a) calcium perchlorate, (b) ammonium sulfate, (c) copper(II) acetate

\*97. (a) sodium carbonate, (b) barium carbonate, (c) nickel(II) acetate

\*98. (a) sodium sulfide, (b) aluminum phosphate, (c) lead(II) acetate

99. (a) Propose a definition for salts, as a class of compounds, on the basis of how they are formed. (b) Provide an example, in the form of a chemical reaction, to illustrate your definition of salts.

100. We can tell from the formula of a salt how it can be produced. Write a balanced chemical equation for the production of each of the following salts: (a) magnesium nitrate, (b) aluminum sulfite, (c) potassium carbonate, (d) zinc chlorate, (e) lithium acetate.

101. Magnesium hydroxide is a gelatinous material that forms during the water purification process in some water treatment plants because of magnesium ions in the water. (a) Write the chemical equation for the reaction of hydrochloric acid with magnesium hydroxide. (b) Explain what drives this reaction to completion.

### Precipitation Reactions

102. A common test for the presence of chloride ions is the formation of a heavy, white precipitate when a solution of silver nitrate is added. (a) Write the balanced chemical equation for the production of silver chloride from silver nitrate solution and calcium chloride solution. (b) Explain why this reaction goes to completion.

103. Based on the solubility guidelines given in Table 4-8, how would you write the formulas for the following substances in a total ionic equation? (a)  $\text{PbSO}_4$ , (b)  $\text{Na}(\text{CH}_3\text{COO})$ , (c)  $\text{Na}_2\text{CO}_3$ , (d)  $\text{MnS}$ , (e)  $\text{BaCl}_2$ .

104. Repeat Exercise 103 for the following: (a)  $(\text{NH}_4)_2\text{SO}_4$ , (b)  $\text{NaBr}$ , (c)  $\text{SrCl}_2$ , (d)  $\text{MgF}_2$ , (e)  $\text{Na}_2\text{CO}_3$ . Refer to the solubility guidelines given in Table 4-8. Classify the compounds in Exercises 105 through 108 as soluble, moderately soluble, or insoluble in water.

105. (a)  $\text{NaClO}_4$ , (b)  $\text{AgCl}$ , (c)  $\text{Pb}(\text{NO}_3)_2$ , (d)  $\text{KOH}$ , (e)  $\text{CaSO}_4$

106. (a)  $\text{BaSO}_4$ , (b)  $\text{Al}(\text{NO}_3)_3$ , (c)  $\text{CuS}$ , (d)  $\text{Na}_3\text{AsO}_4$ , (e)  $\text{Ca}(\text{CH}_3\text{COO})_2$

107. (a)  $\text{Fe}(\text{NO}_3)_3$ , (b)  $\text{Hg}(\text{CH}_3\text{COO})_2$ , (c)  $\text{BeCl}_2$ , (d)  $\text{CuSO}_4$ , (e)  $\text{CaCO}_3$

108. (a)  $\text{KClO}_3$ , (b)  $\text{NH}_4\text{Cl}$ , (c)  $\text{NH}_3$ , (d)  $\text{HNO}_2$ , (e)  $\text{PbS}$

In Exercises 109 and 110, write balanced (i) formula unit, (ii) total ionic, and (iii) net ionic equations for the reactions that occur when *aqueous solutions* of the compounds are mixed.

109. (a)  $\text{Ba}(\text{NO}_3)_2 + \text{K}_2\text{CO}_3 \rightarrow$

(b)  $\text{NaOH} + \text{NiCl}_2 \rightarrow$

(c)  $\text{Al}_2(\text{SO}_4)_3 + \text{NaOH} \rightarrow$

110. (a)  $\text{Cu}(\text{NO}_3)_2 + \text{Na}_2\text{S} \rightarrow$

(b)  $\text{CdSO}_4 + \text{H}_2\text{S} \rightarrow$

(c)  $\text{Bi}_2(\text{SO}_4)_3 + (\text{NH}_4)_2\text{S} \rightarrow$

111. In each of the following, both compounds are water-soluble. Predict whether a precipitate will form when solu-



tions of the two are mixed, and, if so, identify the compound that precipitates. (a)  $\text{Pb}(\text{NO}_3)_2$ ,  $\text{NaI}$ ; (b)  $\text{Ba}(\text{NO}_3)_2$ ,  $\text{KCl}$ ; (c)  $(\text{NH}_4)_2\text{S}$ ,  $\text{AgNO}_3$

112. In each of the following, both compounds are water-soluble. Predict whether a precipitate will form when solutions of the two are mixed, and, if so, identify the compound that precipitates. (a)  $\text{NH}_4\text{Br}$ ,  $\text{Hg}_2(\text{NO}_3)_2$ ; (b)  $\text{KOH}$ ,  $\text{Na}_2\text{S}$ ; (c)  $\text{Cs}_2\text{SO}_4$ ,  $\text{MgCl}_2$

### Identifying Reaction Types

The following reactions apply to Exercises 113 through 121.

- (a)  $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{KOH}(\text{aq}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell)$   
 (b)  $2\text{Rb}(\text{s}) + \text{Br}_2(\ell) \xrightarrow{\text{heat}} 2\text{RbBr}(\text{s})$   
 (c)  $2\text{KI}(\text{aq}) + \text{F}_2(\text{g}) \rightarrow 2\text{KF}(\text{aq}) + \text{I}_2(\text{s})$   
 (d)  $\text{CaO}(\text{s}) + \text{SiO}_2(\text{s}) \xrightarrow{\text{heat}} \text{CaSiO}_3(\text{s})$   
 (e)  $\text{S}(\text{s}) + \text{O}_2(\text{g}) \xrightarrow{\text{heat}} \text{SO}_2(\text{g})$   
 (f)  $\text{BaCO}_3(\text{s}) \xrightarrow{\text{heat}} \text{BaO}(\text{s}) + \text{CO}_2(\text{g})$   
 (g)  $\text{HgS}(\text{s}) + \text{O}_2(\text{g}) \xrightarrow{\text{heat}} \text{Hg}(\ell) + \text{SO}_2(\text{g})$   
 (h)  $\text{AgNO}_3(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{HNO}_3(\text{aq})$   
 (i)  $\text{Pb}(\text{s}) + 2\text{HBr}(\text{aq}) \rightarrow \text{PbBr}_2(\text{s}) + \text{H}_2(\text{g})$   
 (j)  $2\text{HI}(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{I}_2(\text{s}) + 2\text{H}_2\text{O}(\ell)$   
 (k)  $\text{RbOH}(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow \text{RbNO}_3(\text{aq}) + \text{H}_2\text{O}(\ell)$   
 (l)  $\text{N}_2\text{O}_5(\text{s}) + \text{H}_2\text{O}(\ell) \rightarrow 2\text{HNO}_3(\text{aq})$   
 (m)  $\text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g}) \xrightarrow{\text{heat}} \text{H}_2(\text{g}) + \text{CO}_2(\text{g})$   
 (n)  $\text{MgO}(\text{s}) + \text{H}_2\text{O}(\ell) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$   
 (o)  $\text{PbSO}_4(\text{s}) + \text{PbS}(\text{s}) \xrightarrow{\text{heat}} 2\text{Pb}(\text{s}) + 2\text{SO}_2(\text{g})$

113. Identify the precipitation reactions.

114. Identify the acid–base reactions.

115. Identify the oxidation–reduction reactions.

116. Identify the oxidizing agent and reducing agent for each oxidation–reduction reaction.

117. Identify the displacement reactions.

118. Identify the metathesis reactions.

119. Identify the combination reactions.

120. Identify the decomposition reactions.

121. (a) Do any of these reactions fit into more than one class? Which ones? (b) Do any of these reactions not fit into any of our classes of reactions? Which ones?

122. Predict whether or not a solid is formed when we mix the following; identify any solid product by name and identify the reaction type: (a) copper(II) nitrate solution and magnesium metal, (b) barium nitrate and sodium phosphate solutions, (c) calcium acetate solution with aluminum metal, (d) silver nitrate and sodium iodide solutions.

123. Predict whether or not a solid is formed when we mix the following; identify any solid product by formula and by name: (a) potassium permanganate and sodium phosphate solutions, (b) lithium carbonate and cadmium nitrate, (c) stannous fluoride and bismuth chloride, (d) strontium sulfate with barium chloride solutions.

### CONCEPTUAL EXERCISES

124. The following statements are a collection of incorrect statements. Explain why each is wrong. (a) A 1.2 M  $\text{CH}_3\text{COOH}$  solution is a stronger acid than a 0.12 M  $\text{CH}_3\text{COOH}$  solution. (b) The salt produced during the neutralization of nitric acid with potassium hydroxide is  $\text{KNO}_4$ . (c) Nickel will react with a  $\text{HCl}$  solution but not with steam to produce hydrogen and nickel ion solution. Magnesium will react with steam to produce hydrogen and magnesium ions. Therefore, nickel is more reactive than magnesium. As we have seen, two substances may react to form different products when they are mixed in different proportions under different conditions. In Exercises 125 and 126, write balanced equations for the reactions described by words. Assign oxidation numbers.

125. (a) Ethane burns in excess air to form carbon dioxide and water. (b) Ethane burns in a limited amount of air to form carbon monoxide and water. (c) Ethane burns (poorly) in a very limited amount of air to form elemental carbon and water.

126. (a) Butane ( $\text{C}_4\text{H}_{10}$ ) burns in excess air to form carbon dioxide and water. (b) Butane burns in a limited amount of air to form carbon monoxide and water. (c) When heated in the presence of *very little* air, butane “cracks” to form acetylene,  $\text{C}_2\text{H}_2$ , carbon monoxide, and hydrogen.

127. Calcium phosphate is the component of human bone that provides rigidity. Fallout from a nuclear bomb can contain radioactive strontium-90. These two facts are closely tied together when one considers human health. Explain.

128. Limestone consists mainly of the mineral calcite, which is calcium carbonate. A very similar deposit called dolostone is composed primarily of the mineral dolomite, an ionic substance that contains carbonate ions and a mixture of magnesium and calcium ions. (a) Is this a surprising mixture of ions? Explain, based on the periodic table. (b) A test for limestone is to apply cold dilute hydrochloric acid, which causes the rapid formation of bubbles. What causes these bubbles?

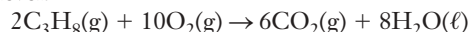


Italian Tourist Bureau

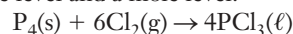
The Dolomite Alps of Italy.

129. Chemical equations can be interpreted on either a particulate level (atoms, molecules, ions) or a mole level (moles of reactants and products). Write word statements to describe

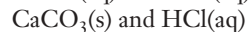
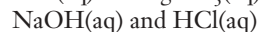
the combustion of propane on a particulate level and a mole level.



- 130.** Write word statements to describe the following reaction on a particulate level and a mole level.



- 131.** When the following pairs of reactants are combined in a beaker: (a) describe in words what the contents of the beaker would look like before and after any reaction that might occur, (b) use different circles for atoms, molecules, and ions to draw a nanoscale (particulate-level) diagram of what the contents would look like, and (c) write a chemical equation for any reactions that might occur.



- 132.** Explain how you could prepare barium sulfate by (a) an acid-base reaction, (b) a precipitation reaction, and (c) a gas-forming reaction. The materials you have to start with are  $\text{BaCO}_3$ ,  $\text{Ba}(\text{OH})_2$ ,  $\text{Na}_2\text{SO}_4$ , and  $\text{H}_2\text{SO}_4$ .

### BUILDING YOUR KNOWLEDGE

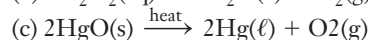
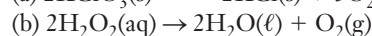
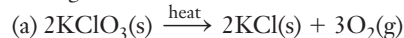
- 133.** Magnesium oxide, marketed as tablets or as an aqueous slurry called “milk of magnesia,” is a common commercial antacid. What volume, in milliliters, of fresh gastric juice, corresponding in acidity to  $0.17\text{ M HCl}$ , could be neutralized by  $104\text{ mg}$  of magnesium oxide?



Charles D. Winters

Milk of Magnesia

- 134.** How many moles of oxygen can be obtained by the decomposition of  $15.0\text{ grams}$  of reactant in each of the following reactions?



### BEYOND THE TEXTBOOK

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