CHAPTER 8

Chemical Equations and Reactions

The evolution of light and heat is an indication that a chemical reaction is taking place.

Describing Chemical Reactions

SECTION 8-1

Objectives

- List three observations that suggest that a chemical reaction has taken place.
- List three requirements for a correctly written chemical equation.
- Write a word equation and a formula equation for a given chemical reaction.
- Balance a formula equation by inspection.

A *chemical reaction* is the process by which one or more substances are changed into one or more different substances. In any chemical reaction, the original substances are known as the *reactants* and the resulting substances are known as the *products*. According to the law of conservation of mass, the total mass of reactants must equal the total mass of products for any given chemical reaction.

Chemical reactions are described by chemical equations. *A* **chemical equation** *represents, with symbols and formulas, the identities and relative amounts of the reactants and products in a chemical reaction.* For example, the following chemical equation shows that the reactant ammonium dichromate yields the products nitrogen, chromium(III) oxide, and water.

$$(NH_4)_2Cr_2O_7(s) \longrightarrow N_2(g) + Cr_2O_3(s) + 4H_2O(g)$$

This strongly exothermic reaction is shown in Figure 8-1.

Indications of a Chemical Reaction

To know for certain that a chemical reaction has taken place requires evidence that one or more substances have undergone a change in identity. Absolute proof of such a change can be provided only by chemical analysis of the products. However, certain easily observed changes usually indicate that a chemical reaction has occurred.

1. *Evolution of heat and light.* A change in matter that releases energy as both heat and light is strong evidence that a chemical reaction has taken place. For example, you can see in Figure 8-1 that the decomposition of ammonium dichromate is accompanied by the evolution of much heat and light. And you can see evidence that a chemical reaction occurs between natural gas and oxygen if you burn gas for cooking in your house. Some reactions release only heat or only light. But the evolution of heat or light by itself is not necessarily a sign of chemical change because many physical changes also release either heat or light.

FIGURE 8-1 The decomposition of ammonium dichromate proceeds rapidly, releasing energy in the form of light and heat.

FIGURE 8-2 (a) The reaction of vinegar and baking soda is evidenced by the production of bubbles of carbon dioxide gas. (b) When water solutions of ammonium sulfide and cadmium nitrate are combined, the yellow precipitate cadmium sulfide forms.

(a) (b)

2. *Production of a gas.* The evolution of gas bubbles when two substances are mixed is often evidence of a chemical reaction. For example, bubbles of carbon dioxide gas form immediately when baking soda is mixed with vinegar, in the vigorous reaction that is shown in Figure 8-2(a).
3. *Formation of a precipitate.* Many chemical reactions take place between substances that are dissolved in liquids. If a solid appears after two solutions are mixed, a reaction has likely occurred. *A solid that is produced as a result of a chemical reaction in solution and that separates from the solution is known as a* **precipitate.** A precipitate-forming reaction is shown in Figure 8-2(b).
4. *Color change.* A change in color is often an indication of a chemical reaction.

internet**connect**

SCILINKS NSTA

TOPIC: Chemical reactions
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2081

Characteristics of Chemical Equations

A properly written chemical equation can summarize any chemical change. The following requirements will aid you in writing and reading chemical equations correctly.

1. *The equation must represent known facts.* All reactants and products must be identified, either through chemical analysis in the laboratory or from sources that give the results of experiments.
2. *The equation must contain the correct formulas for the reactants and products.* Remember what you learned in Chapter 7 about symbols and formulas. Knowledge of the common oxidation states of the elements and of methods of writing formulas will enable you to supply formulas for reactants and products if they are not available. Recall that the elements listed in Table 8-1 exist primarily as diatomic molecules, such as H_2 and O_2. Each of these elements is represented in an equation by its molecular formula. Other elements in the elemental state are usually represented simply by their atomic symbols. For example, iron is represented as Fe and carbon is represented as C. The symbols are not given any subscripts because the elements do not

TABLE 8-1 *Elements That Normally Exist as Diatomic Molecules*

Element	Symbol	Molecular formula	Physical state at room temperature
Hydrogen	H	H_2	gas
Nitrogen	N	N_2	gas
Oxygen	O	O_2	gas
Fluorine	F	F_2	gas
Chlorine	Cl	Cl_2	gas
Bromine	Br	Br_2	liquid
Iodine	I	I_2	solid

form definite molecular structures. Two exceptions to this rule are sulfur, which is usually written S_8, and phosphorus, which is usually written P_4. In these cases, the formulas reflect each element's unique atomic arrangement in its natural state.

3. *The law of conservation of mass must be satisfied.* Atoms are neither created nor destroyed in ordinary chemical reactions. Therefore, the same number of atoms of each element must appear on each side of a correct chemical equation. To equalize numbers of atoms, coefficients are added where necessary. *A* **coefficient** *is a small whole number that appears in front of a formula in a chemical equation.* Placing a coefficient in front of a formula specifies the relative number of moles of the substance; if no coefficient is written, the coefficient is assumed to be 1. For example, the coefficient 4 in the equation on page 241 indicates that 4 mol of water are produced for each mole of nitrogen and chromium(III) oxide that is produced.

Word and Formula Equations

The first step in writing a chemical equation is to identify the facts to be represented. It is often helpful to write a **word equation,** *an equation in which the reactants and products in a chemical reaction are represented by words.* A word equation has only qualitative (descriptive) meaning. It does not give the whole story because it does not give the quantities of reactants used or products formed.

Consider the reaction of methane, the principal component of natural gas, with oxygen. When methane burns in air, it combines with oxygen to produce carbon dioxide and water vapor. In the reaction, methane and oxygen are the reactants, and carbon dioxide and water are the products. The word equation for the reaction of methane and oxygen is written as follows.

$$\text{methane} + \text{oxygen} \longrightarrow \text{carbon dioxide} + \text{water}$$

The arrow, $\longrightarrow$, is read as *react to yield* or *yield* (also *produce* or *form*). So the equation above is read, "methane and oxygen react to yield carbon dioxide and water," or simply, "methane and oxygen yield carbon dioxide and water."

The next step in writing a correct chemical equation is to replace the names of the reactants and products with appropriate symbols and formulas. Methane is a molecular compound composed of one carbon atom and four hydrogen atoms. Its chemical formula is CH_4. Recall that oxygen exists in nature as diatomic molecules; it is therefore represented as O_2. The correct formulas for carbon dioxide and water are CO_2 and H_2O, respectively.

A **formula equation** *represents the reactants and products of a chemical reaction by their symbols or formulas.* The formula equation for the reaction of methane and oxygen is written as follows.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g) \quad \text{(not balanced)}$$

The *g* in parentheses after each formula indicates that the corresponding substance is in the gaseous state. Like a word equation, a formula equation is a qualitative statement. It gives no information about the amounts of reactants or products.

A formula equation meets two of the three requirements for a correct chemical equation. It represents the facts and shows the correct symbols and formulas for the reactants and products. To complete the process of writing a correct equation, the law of conservation of mass must be taken into account. The relative amounts of reactants and products represented in the equation must be adjusted so that the numbers and types of atoms are the same on both sides of the equation. This process is called *balancing an equation* and is carried out by inserting coefficients. Once it is balanced, a formula equation is a correctly written chemical equation.

Look again at the formula equation for the reaction of methane and oxygen.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g) \quad \text{(not balanced)}$$

CHEMISTRY INTERACTIVE TUTOR

Module 5: Equations and Stoichiometry

To balance the equation, begin by counting atoms of elements that are combined with atoms of other elements and that appear only once on each side of the equation. In this case, we could begin by counting either carbon or hydrogen atoms. Usually, the elements hydrogen and oxygen are balanced only after balancing all other elements in an equation. (You will read more about the rules of balancing equations later in the chapter.) Thus, we begin by counting carbon atoms.

Inspecting the formula equation reveals that there is one carbon atom on each side of the arrow. Therefore, carbon is already balanced in the equation. Counting hydrogen atoms reveals that there are four hydrogen atoms in the reactants but only two in the products. Two additional hydrogen atoms are needed on the right side of the equation. They can be added by placing the coefficient 2 in front of the chemical formula H_2O.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) \quad \text{(partially balanced)}$$

A coefficient multiplies the number of atoms of each element indicated in a chemical formula. Thus, $2H_2O$ represents *four* H atoms and *two* O atoms. To add two more hydrogen atoms to the right side of the equation, one may be tempted to change the subscript in the formula of water so that H_2O becomes H_4O. However, this would be a mistake because changing the subscripts of a chemical formula changes the *identity* of the compound. H_4O is not a product in the combustion of methane. In fact, there is no such compound. One must use only coefficients to change the relative number of atoms in a chemical equation because coefficients change the numbers of atoms without changing the identities of the reactants or products.

Now consider the number of oxygen atoms. There are four oxygen atoms on the right side of the arrow in the partially balanced equation. Yet there are only two oxygen atoms on the left side of the arrow. One can increase the number of oxygen atoms on the left side to four by placing the coefficient 2 in front of the molecular formula for oxygen. This results in a correct chemical equation, or *balanced formula equation,* for the burning of methane in oxygen.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

This reaction is further illustrated in Figure 8-3.

Additional Symbols Used in Chemical Equations

Table 8-2 on page 246 summarizes the symbols commonly used in chemical equations. Sometimes a gaseous product is indicated by an arrow pointing upward, ↑, instead of (g), as shown in the table. A downward arrow, ↓, is often used to show the formation of a precipitate during a reaction in solution.

The conditions under which a reaction takes place are often indicated by placing information above or below the reaction arrow. The word *heat,*

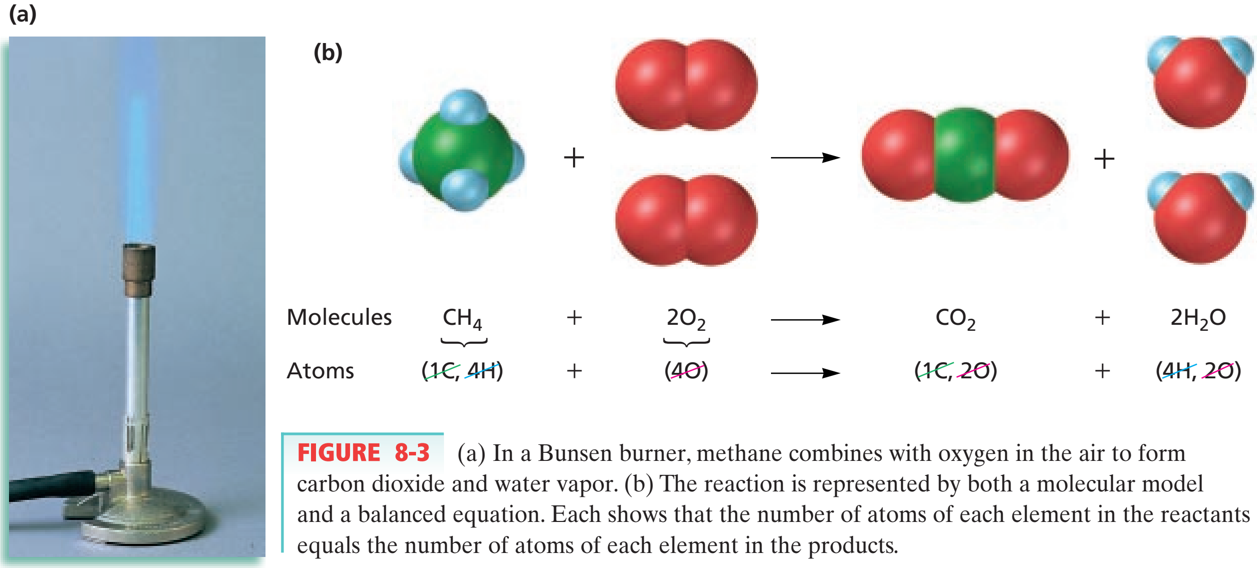


FIGURE 8-3 (a) In a Bunsen burner, methane combines with oxygen in the air to form carbon dioxide and water vapor. (b) The reaction is represented by both a molecular model and a balanced equation. Each shows that the number of atoms of each element in the reactants equals the number of atoms of each element in the products.

TABLE 8-2 Symbols Used in Chemical Equations

Symbol	Explanation
$\longrightarrow$	"Yields"; indicates result of reaction
$\rightleftarrows$	Used in place of a single arrow to indicate a reversible reaction
(s)	A reactant or product in the solid state; also used to indicate a precipitate
$\downarrow$	Alternative to (s), but used only to indicate a precipitate
(l)	A reactant or product in the liquid state
(aq)	A reactant or product in an aqueous solution (dissolved in water)
(g)	A reactant or product in the gaseous state
$\uparrow$	Alternative to (g), but used only to indicate a gaseous product
$\xrightarrow{\Delta}$ or $\xrightarrow{\text{heat}}$	Reactants are heated
$\xrightarrow{\text{2 atm}}$	Pressure at which reaction is carried out, in this case 2 atm
$\xrightarrow{\text{pressure}}$	Pressure at which reaction is carried out exceeds normal atmospheric pressure
$\xrightarrow{0°\text{C}}$	Temperature at which reaction is carried out, in this case 0°C
$\xrightarrow{MnO_2}$	Formula of catalyst, in this case manganese dioxide, used to alter the rate of the reaction

symbolized by a Greek capital delta, Δ, indicates that the reactants must be heated. The specific temperature at which a reaction occurs may also be written over the arrow. For some reactions, it is important to specify the pressure at which the reaction occurs or to specify that the pressure must be above normal. Many reactions are speeded up and can take place at lower temperatures in the presence of a *catalyst.* A catalyst is a substance that changes the rate of a chemical reaction but can be recovered unchanged. To show that a catalyst is present, the formula for the catalyst or the word *catalyst* is written over the reaction arrow.

In many reactions, as soon as the products begin to form, they immediately begin to react with each other and re-form the reactants. In other words, the reverse reaction also occurs. The reverse reaction may occur to a greater or lesser degree than the original reaction, depending on the specific reaction and the conditions. *A* **reversible reaction** *is a chemical reaction in which the products re-form the original* *reactants.* The reversibility of a reaction is indicated by writing two arrows pointing in opposite directions. For example, the reversible reaction between iron and water vapor is written as follows.

$$3Fe(s) + 4H_2O(g) \rightleftarrows Fe_3O_4(s) + 4H_2(g)$$

With an understanding of all the symbols and formulas used, it is possible to translate a chemical equation into a sentence. Consider the following equation.

$$2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$$

Translated into a sentence, this equation reads, "When heated, solid mercury(II) oxide yields liquid mercury and gaseous oxygen."

It is also possible to write a chemical equation from a sentence describing a reaction. Consider the sentence, "Under pressure and in the presence of a platinum catalyst, gaseous ethene and hydrogen form gaseous ethane." This sentence can be translated into the following equation.

$$C_2H_4(g) + H_2(g) \xrightarrow[\text{Pt}]{\text{pressure}} C_2H_6(g)$$

Throughout this chapter we will often include the symbols for physical states (*s, l, g,* and *aq*) in balanced formula equations. You should be able to interpret these symbols when they are used and to supply them when the necessary information is available.

SAMPLE PROBLEM 8-1

Write word and formula equations for the chemical reaction that occurs when solid sodium oxide is added to water at room temperature and forms sodium hydroxide (dissolved in the water). Include symbols for physical states in the formula equation. Then balance the formula equation to give a balanced chemical equation.

SOLUTION The word equation must show the reactants, sodium oxide and water, to the left of the arrow. The product, sodium hydroxide, must appear to the right of the arrow.

$$\text{sodium oxide} + \text{water} \longrightarrow \text{sodium hydroxide}$$

The word equation is converted to a formula equation by replacing the name of each compound with the appropriate chemical formula. To do this requires knowing that sodium has an oxidation state of +1, that oxygen usually has an oxidation state of –2, and that a hydroxide ion has a charge of 1–.

$$Na_2O + H_2O \longrightarrow NaOH \quad \text{(not balanced)}$$

Adding symbols for the physical states of the reactants and products and the coefficient 2 in front of NaOH produces a balanced chemical equation.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

SAMPLE PROBLEM 8-2

Translate the following chemical equation into a sentence:

$$PbCl_2(aq) + Na_2CrO_4(aq) \longrightarrow PbCrO_4(s) + 2NaCl(aq)$$

SOLUTION Each reactant is an ionic compound and is named according to the rules for such compounds. Both reactants are in aqueous solution. One product is a precipitate and the other remains in solution. The equation is translated as follows: Aqueous solutions of lead(II) chloride and sodium chromate react to produce a precipitate of lead(II) chromate plus sodium chloride in aqueous solution.

PRACTICE

1. Write word and balanced chemical equations for the following reactions. Include symbols for physical states when indicated.
   a. Solid calcium reacts with solid sulfur to produce solid calcium sulfide.
   b. Hydrogen gas reacts with fluorine gas to produce hydrogen fluoride gas. (Hint: See Table 8-1.)
   c. Solid aluminum metal reacts with aqueous zinc chloride to produce solid zinc metal and aqueous aluminum chloride.

Answer

1. a. calcium + sulfur $\longrightarrow$ calcium sulfide; $8Ca(s) + S_8(s) \longrightarrow 8CaS(s)$
   b. hydrogen + fluorine $\longrightarrow$ hydrogen fluoride; $H_2(g) + F_2(g) \longrightarrow 2HF(g)$
   c. aluminum + zinc chloride $\longrightarrow$ zinc + aluminum chloride; $2Al(s) + 3ZnCl_2(aq) \longrightarrow 3Zn(s) + 2AlCl_3(aq)$

2. Translate the following chemical equations into sentences:
   a. $CS_2(l) + 3O_2(g) \longrightarrow CO_2(g) + 2SO_2(g)$
   b. $NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$

2. a. Liquid carbon disulfide reacts with oxygen gas to produce carbon dioxide gas and sulfur dioxide gas.
   b. Aqueous solutions of sodium chloride and silver nitrate react to produce aqueous sodium nitrate and a precipitate of silver chloride.

Significance of a Chemical Equation

Chemical equations are very useful in doing quantitative chemical work. The arrow in a balanced chemical equation is like an equal sign. And the chemical equation as a whole is similar to an algebraic equation in that it expresses an equality. Let's examine some of the quantitative information revealed by a chemical equation.

1. *The coefficients of a chemical reaction indicate relative, not absolute, amounts of reactants and products.* A chemical equation usually shows the smallest numbers of atoms, molecules, or ions that will satisfy the law of conservation of mass in a given chemical reaction.

Consider the equation for the formation of hydrogen chloride from hydrogen and chlorine.

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

The equation indicates that 1 molecule of hydrogen reacts with 1 molecule of chlorine to produce 2 molecules of hydrogen chloride, giving the following molecular ratio of reactants and products.

$$1 \text{ molecule } H_2 : 1 \text{ molecule } Cl_2 : 2 \text{ molecules HCl}$$

This ratio shows the smallest possible relative amounts of the reaction's reactants and products. To obtain larger relative amounts, we simply multiply each coefficient by the same number. Thus, 20 molecules of hydrogen would react with 20 molecules of chlorine to yield 40 molecules of hydrogen chloride. The reaction can also be considered in terms of amounts in moles: 1 mol of hydrogen molecules reacts with 1 mol of chlorine molecules to yield 2 mol of hydrogen chloride molecules.

2. *The relative masses of the reactants and products of a chemical reaction can be determined from the reaction's coefficients.* Recall from Figure 7-4 on page 224 that an amount of an element or compound in moles can be converted to a mass in grams by multiplying by the appropriate molar mass. We know that 1 mol of hydrogen reacts with 1 mol of chlorine to yield 2 mol of hydrogen chloride. The relative masses of the reactants and products are calculated as follows.

CHEMISTRY INTERACTIVE TUTOR

Module 5: Equations and Stoichiometry

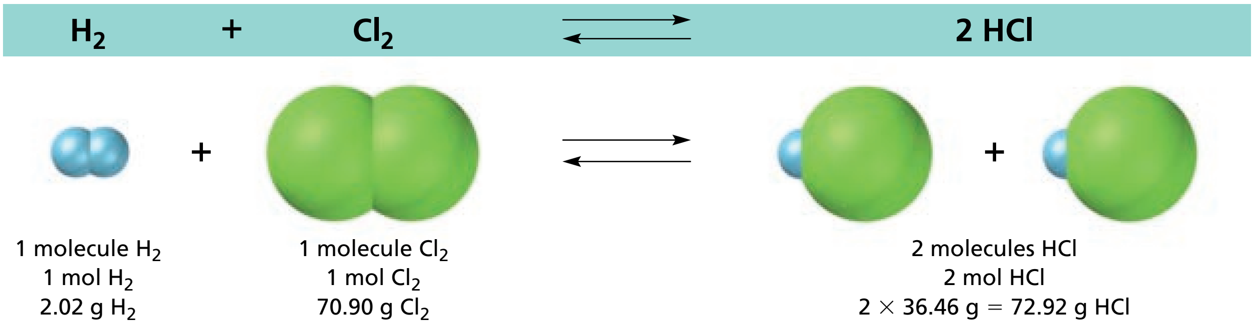
$$\cancel{1 \text{ mol } H_2} \times \frac{2.02 \text{ g } H_2}{\cancel{\text{mol } H_2}} = 2.02 \text{ g } H_2$$

$$\cancel{1 \text{ mol } Cl_2} \times \frac{70.90 \text{ g } Cl_2}{\cancel{\text{mol } Cl_2}} = 70.90 \text{ g } Cl_2$$

$$\cancel{2 \text{ mol HCl}} \times \frac{36.46 \text{ g HCl}}{\cancel{\text{mol HCl}}} = 72.92 \text{ g HCl}$$

The chemical equation shows that 2.02 g of hydrogen will react with 70.90 g of chlorine to yield 72.92 g of hydrogen chloride.

FIGURE 8-4 This representation of the reaction of hydrogen and chlorine to yield hydrogen chloride shows several ways to interpret the quantitative information of a chemical reaction.



3. *The reverse reaction for a chemical equation has the same relative amounts of substances as the forward reaction.* Because a chemical equation is like an algebraic equation, the equality can be read in either direction. Reading the hydrogen chloride formation equation on page 249 from right to left, we can see that 2 molecules of hydrogen chloride break down to form 1 molecule of hydrogen plus 1 molecule of chlorine. Similarly, 2 mol (72.92 g) of hydrogen chloride yield 1 mol (2.02 g) of hydrogen and 1 mol (70.90 g) of chlorine.

We have seen that a chemical equation provides useful quantitative information about a chemical reaction. However, there is also important information that is *not* provided by a chemical equation. For instance, an equation gives no indication of whether a reaction will actually occur. A chemical equation can be written for a reaction that may not even take place. Some guidelines about the types of simple reactions that can be expected to occur are given in Sections 8-2 and 8-3. And later chapters provide additional guidelines for other types of reactions. In all these guidelines, it is important to remember that experimentation forms the basis for confirming that a particular chemical reaction will occur.

In addition, chemical equations give no information about the speed at which reactions occur or about how the bonding between atoms or ions changes during the reaction. These aspects of chemical reactions are discussed in Chapter 17.

FIGURE 8-5 When an electric current is passed through water that has been made slightly conductive, the water molecules break down to yield hydrogen (in tube at right) and oxygen (in tube at left). Bubbles of each gas are evidence of the reaction. Note that twice as much hydrogen as oxygen is produced.

Balancing Chemical Equations

Most of the equations in the remainder of this chapter can be balanced by inspection. The following procedure demonstrates how to master balancing equations by inspection using a step-by-step approach. The equation for the decomposition of water (see Figure 8-5) will be used as an example.

1. *Identify the names of the reactants and the products, and write a word equation.* The word equation for the reaction shown in Figure 8-5 is written as follows.

$$\text{water} \longrightarrow \text{hydrogen} + \text{oxygen}$$

2. *Write a formula equation by substituting correct formulas for the names of the reactants and the products.* We know that the formula for water is H_2O. And recall that both hydrogen and oxygen exist as diatomic molecules. Therefore, their correct formulas are H_2 and O_2, respectively.

$$H_2O(l) \longrightarrow H_2(g) + O_2(g) \quad \text{(not balanced)}$$

3. *Balance the formula equation according to the law of conservation of mass.* This last step is done by trial and error. Coefficients are changed and the numbers of atoms are counted on both sides of the equation. When the numbers of each type of atom are the same for both the products and the reactants, the equation is balanced. The trial-and-error method of balancing equations is made easier by the use of the following guidelines.

   - Balance the different types of atoms one at a time.
   - First balance the atoms of elements that are combined and that appear only once on each side of the equation.
   - Balance polyatomic ions that appear on both sides of the equation as single units.
   - Balance H atoms and O atoms after atoms of all other elements have been balanced.

The formula equation in our example shows that there are two oxygen atoms on the right and only one on the left. To balance oxygen atoms, the number of H_2O molecules must be increased. Placing the coefficient 2 before H_2O gives the necessary two oxygen atoms on the left.

$$2H_2O(l) \longrightarrow H_2(g) + O_2(g) \quad \text{(partially balanced)}$$

The coefficient 2 in front of H_2O has upset the balance of hydrogen atoms. Placing the coefficient 2 in front of hydrogen, H_2, on the right, gives an equal number of hydrogen atoms (4) on both sides of the equation.

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

4. *Count atoms to be sure that the equation is balanced.* Make sure that equal numbers of atoms of each element appear on both sides of the arrow.

$$\begin{array}{ccccc} 2H_2O(l) & \longrightarrow & 2H_2(g) & + & O_2(g) \\ (\cancel{4H} + \cancel{2O}) & = & (\cancel{4H}) & + & (\cancel{2O}) \end{array}$$

Occasionally at this point, the coefficients do not represent the smallest possible whole-number ratio of reactants and products. When this happens, the coefficients should be divided by their greatest common factor in order to obtain the smallest possible whole-number coefficients.

Balancing chemical equations by inspection becomes easier as you gain experience. Learn to avoid the most common mistakes: (1) writing incorrect chemical formulas for reactants or products and (2) trying to balance an equation by changing subscripts. Remember that subscripts cannot be added, deleted, or changed. Eventually, you will probably be able to skip writing the word equation and each separate step. However, *do not* leave out the final step of counting atoms to be sure the equation is balanced.

SAMPLE PROBLEM 8-3

The reaction of zinc with aqueous hydrochloric acid produces a solution of zinc chloride and hydrogen gas. This reaction is shown at right in Figure 8-6. Write a balanced chemical equation for the reaction.

FIGURE 8-6 Solid zinc reacts with hydrochloric acid to form aqueous zinc chloride and hydrogen gas.

SOLUTION

1 ANALYZE *Write the word equation.*

$$\text{zinc} + \text{hydrochloric acid} \longrightarrow \text{zinc chloride} + \text{hydrogen}$$

2 PLAN *Write the formula equation.*

$$Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g) \quad \text{(not balanced)}$$

3 COMPUTE *Adjust the coefficients.* Note that chlorine and hydrogen each appear only once on each side of the equation. We balance chlorine first because it is combined on both sides of the equation. Also, recall from the guidelines on page 251 that hydrogen and oxygen are balanced only after all other elements in the reaction are balanced. To balance chlorine, we place the coefficient 2 before HCl. Two molecules of hydrogen chloride also yield the required two hydrogen atoms on the right. Finally, note that there is one zinc atom on each side in the formula equation. Therefore, no further coefficients are needed.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

4 EVALUATE *Count atoms to check balance.*

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$
$$(1\cancel{Zn}) + (2\cancel{H} + 2\cancel{Cl}) = (1\cancel{Zn} + 2\cancel{Cl}) + (2\cancel{H})$$

The equation is balanced.

PRACTICE

1. Write word, formula, and balanced chemical equations for each of the following reactions:
   a. Magnesium and hydrochloric acid react to produce magnesium chloride and hydrogen.
   b. Aqueous nitric acid reacts with solid magnesium hydroxide to produce aqueous magnesium nitrate and water.

Answer

1. a. Word: magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen
   Formula: $Mg + HCl \longrightarrow MgCl_2 + H_2$
   Balanced: $Mg + 2HCl \longrightarrow MgCl_2 + H_2$

   b. Word: nitric acid + magnesium hydroxide $\longrightarrow$ magnesium nitrate + water
   Formula: $HNO_3(aq) + Mg(OH)_2(s) \longrightarrow Mg(NO_3)_2(aq) + H_2O(l)$
   Balanced: $2HNO_3(aq) + Mg(OH)_2(s) \longrightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$

SAMPLE PROBLEM 8-4

Solid aluminum carbide, Al_4C_3, reacts with water to produce methane gas and solid aluminum hydroxide. Write a balanced chemical equation for this reaction.

SOLUTION

The reactants are aluminum carbide and water. The products are methane and aluminum hydroxide. The formula equation is written as follows.

$$Al_4C_3(s) + H_2O(l) \longrightarrow CH_4(g) + Al(OH)_3(s) \quad \text{(not balanced)}$$

Begin balancing the formula equation by counting either aluminum atoms or carbon atoms. (Remember that hydrogen and oxygen atoms are balanced last.) There are four Al atoms on the left. To balance Al atoms, place the coefficient 4 before $Al(OH)_3$ on the right.

$$Al_4C_3(s) + H_2O(l) \longrightarrow CH_4(g) + 4Al(OH)_3(s) \quad \text{(partially balanced)}$$

Now balance the carbon atoms. With three C atoms on the left, the coefficient 3 must be placed before CH_4 on the right.

$$Al_4C_3(s) + H_2O(l) \longrightarrow 3CH_4(g) + 4Al(OH)_3(s) \quad \text{(partially balanced)}$$

Balance oxygen atoms next because oxygen, unlike hydrogen, appears only once on each side of the equation. There is one O atom on the left and 12 O atoms in the four $Al(OH)_3$ formula units on the right. Placing the coefficient 12 before H_2O balances the O atoms.

$$Al_4C_3(s) + 12H_2O(l) \longrightarrow 3CH_4(g) + 4Al(OH)_3(s)$$

This leaves the hydrogen atoms to be balanced. There are 24 H atoms on the left. On the right, there are 12 H atoms in the methane molecules and 12 in the aluminum hydroxide formula units, totaling 24 H atoms. The H atoms are balanced.

$$\begin{array}{ccccccc} Al_4C_3(s) & + & 12H_2O(l) & \longrightarrow & 3CH_4(g) & + & 4Al(OH)_3(s) \\ (\cancel{4Al} + \cancel{3C}) & + & (\cancel{24H} + \cancel{12O}) & = & (\cancel{3C} + \cancel{12H}) & + & (\cancel{4Al} + \cancel{12H} + \cancel{12O}) \end{array}$$

The equation is balanced.

SAMPLE PROBLEM 8-5

Aluminum sulfate and calcium hydroxide are used in a water-purification process. When added to water, they dissolve and react to produce two insoluble products, aluminum hydroxide and calcium sulfate. These products settle out, taking suspended solid impurities with them. Write a balanced chemical equation for the reaction.

SOLUTION

Each of the reactants and products is an ionic compound. Recall from Chapter 7 that the formulas of ionic compounds are determined by the charges of the ions composing each compound. The formula reaction is thus written as follows.

$$Al_2(SO_4)_3 + Ca(OH)_2 \longrightarrow Al(OH)_3 + CaSO_4 \quad \text{(not balanced)}$$

There is one Ca atom on each side of the equation, so the calcium atoms are already balanced. There are two Al atoms on the left and one Al atom on the right. Placing the coefficient 2 in front of $Al(OH)_3$ produces the same number of Al atoms on each side of the equation.

$$Al_2(SO_4)_3 + Ca(OH)_2 \longrightarrow 2Al(OH)_3 + CaSO_4 \text{ (partially balanced)}$$

Next, checking SO_4^{2-} ions shows that there are three SO_4^{2-} ions on the left side of the equation and only one on the right side. Placing the coefficient 3 before $CaSO_4$ gives an equal number of SO_4^{2-} ions on each side.

$$Al_2(SO_4)_3 + Ca(OH)_2 \longrightarrow 2Al(OH)_3 + 3CaSO_4 \text{ (partially balanced)}$$

There are now three Ca atoms on the right, however. By placing the coefficient 3 in front of $Ca(OH)_2$, we once again have an equal number of Ca atoms on each side. This last step also gives six OH^- ions on both sides of the equation.

$$Al_2(SO_4)_3(aq) + 3Ca(OH)_2(aq) \longrightarrow 2Al(OH)_3(s) + 3CaSO_4(s)$$
$$(\cancel{2Al} + \cancel{3SO_4^{2-}}) + (\cancel{3Ca} + \cancel{6OH^-}) = (\cancel{2Al} + \cancel{6OH^-}) + (\cancel{3Ca} + \cancel{3SO_4^{2-}})$$

The equation is balanced.

PRACTICE

1. Write balanced chemical equations for each of the following reactions:
   a. Solid sodium combines with chlorine gas to produce solid sodium chloride.
   b. When solid copper reacts with aqueous silver nitrate, the products are aqueous copper(II) nitrate and solid silver.
   c. In a blast furnace, the reaction between solid iron(III) oxide and carbon monoxide gas produces solid iron and carbon dioxide gas.

Answer

1. a. $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

b. $Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

c. $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$

SECTION REVIEW

1. Describe the differences between word equations, formula equations, and chemical equations.
2. Write word and formula equations for the reaction in which aqueous solutions of sulfuric acid and sodium hydroxide react to form aqueous sodium sulfate and water.
3. Translate the following chemical equations into sentences:
   a. $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
   b. $2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$
4. Write the word, formula, and chemical equations for the reaction between hydrogen sulfide gas and oxygen gas that produces sulfur dioxide gas and water vapor.
5. Write the chemical equation for each of the following reactions:
   a. ammonium chloride + calcium hydroxide $\longrightarrow$ calcium chloride + ammonia + water
   b. hexane, C_6H_{14}, + oxygen $\longrightarrow$ carbon dioxide + water

A Chemical Mystery

From "The Chemical Adventures of Sherlock Holmes: The Hound of Henry Armitage" by Thomas G. Waddell and Thomas R. Rybolt in *The Journal of Chemical Education*

"I knew it," the old man snapped. "He was poisoned, wasn't he? . . ."

. . . But Holmes was not listening. He had picked up the dog's bowl, now empty, and was vigorously sniffing, not unlike the hound itself, at the crusted remains of the last meal . . .

An hour later I was in my chair at 221B Baker Street. Holmes was in his laboratory and I could hear him humming. In the background was the usual clattering and clanking of laboratory equipment . . . Suddenly, Holmes called to me.

"Watson, come here. I need you." . . . He calmly scribbled an equation on a slip of paper and handed it to me. "If you can balance this equation, Watson, you can solve this mystery." I looked at the page as best I could and saw the following equation with the formula of a reactant clearly missing.

$$C_6H_5NH_2 + 3KOH + \_\_\_\_ \Rightarrow C_6H_5NC + 3KCl + 3H_2O$$

Holmes paced back and forth with his hands clasped behind his back. "One part aniline, three parts potassium hydroxide, and one part unknown poison yields one part phenylisocyanide, three parts potassium chloride, and three parts water. The missing reactant can be identified by balancing the equation with respect to all the atoms involved. The product phenylisocyanide . . . is derived by this reaction from that missing chemical which was the poison deliberately placed in the hound's food."

"I can follow you part of the way," I submitted. "You undoubtedly detected a foreign substance in the dog food due to a characteristic aroma."

"Correct, Watson," Holmes replied. "And as a chemist I knew immediately that the poison was *volatile* . . . We observed the compound to be a liquid at room temperature, immiscible with water, and having a density greater than 1.00 g/mL! The unpleasant sweetness of it was also very helpful. The possibilities were quite limited at that point, Watson . . .

I formed a working hypothesis and performed a known chemical test for such a poisonous liquid meeting all these criteria. Did you balance the equation, Watson? The equation confirms it!"

"I can do it, Holmes. I remember that much chemistry. Let me see . . . the missing reactant must have chlorine . . . 3 units to balance Cl in the product!"

"Very good, Watson. Go on with it."

"It gets more complex, now, but look, there is one extra carbon atom in the products! Is CCl_3 the compound?"

"Carbon makes *four* bonds, Watson, not three," said Holmes with a frown.

"I have it! $CHCl_3$ balances the equation! That's *chloroform,* Holmes! Of course. It all is consistent."

Reading for Meaning

Can you infer the meaning of the word *volatile* from the story? Write down your definition. Then compare your definition with one from a chemical or technical dictionary.

internetconnect

SCILINKS NSTA

TOPIC: Chloroform
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2082

SECTION 8-2

Types of Chemical Reactions

OBJECTIVES

- Define and give general equations for *synthesis, decomposition, single-replacement,* and *double-replacement* reactions.
- Classify a reaction as synthesis, decomposition, single-replacement, double-replacement, or combustion.
- List three types of synthesis reactions and six types of decomposition reactions.
- List four types of single-replacement reactions and three types of double-replacement reactions.
- Predict the products of simple reactions given the reactants.

Thousands of known chemical reactions occur in living systems, in industrial processes, and in chemical laboratories. Often it is necessary to predict the products formed in one of these reactions. Memorizing the equations for so many chemical reactions would be a difficult task. It is therefore more useful and realistic to classify reactions according to various similarities and regularities. This general information about reaction types can then be used to predict the products of specific reactions.

There are several different ways to classify chemical reactions, and none are entirely satisfactory. The classification scheme described in this section provides an introduction to five basic types of reactions: synthesis, decomposition, single-replacement, double-replacement, and combustion. In later chapters you will be introduced to categories that are useful in classifying other types of chemical reactions.

Synthesis Reactions

In a **synthesis reaction,** *also known as a* **composition reaction,** *two or more substances combine to form a new compound.* This type of reaction is represented by the following general equation.

$$A + X \longrightarrow AX$$

A and X can be elements or compounds. AX is a compound. The following examples illustrate several kinds of synthesis reactions.

Reactions of Elements with Oxygen and Sulfur

One simple type of synthesis reaction is the combination of an element with oxygen to produce an *oxide* of the element. Almost all metals react with oxygen to form oxides. For example, when a thin strip of magnesium metal is placed in an open flame, it burns with bright white light. When the metal strip is completely burned, only a fine white powder of magnesium oxide is left. This chemical reaction, shown in Figure 8-7, is represented by the following equation.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

The other Group 2 elements react in a similar manner, forming oxides with the formula MO, where M represents the metal. The Group 1 metals form oxides with the formula M_2O, for example, Li_2O. The Group 1 and Group 2 elements react similarly with sulfur, forming *sulfides* with the formulas M_2S and MS, respectively. Examples of these types of synthesis reactions are shown below.

$$16Rb(s) + S_8(s) \longrightarrow 8Rb_2S(s)$$
$$8Ba(s) + S_8(s) \longrightarrow 8BaS(s)$$

Some metals, such as iron, combine with oxygen to produce two different oxides.

$$2Fe(s) + O_2(g) \longrightarrow 2FeO(s)$$
$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$$

In the product of the first reaction, iron is in an oxidation state of +2. In the product of the second reaction, iron is in an oxidation state of +3. The particular oxide formed depends on the conditions surrounding the reactants. Both oxides are shown below in Figure 8-8.

Nonmetals also undergo synthesis reactions with oxygen to form oxides. Sulfur, for example, reacts with oxygen to form sulfur dioxide. And when carbon is burned in air, carbon dioxide is produced.

$$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$
$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

In a limited supply of oxygen, carbon monoxide is formed.

$$2C(s) + O_2(g) \longrightarrow 2CO(g)$$

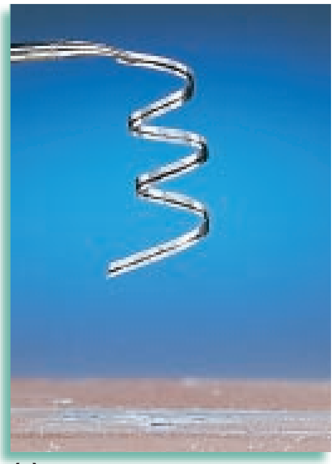
Hydrogen reacts with oxygen to form dihydrogen monoxide, better known as water.

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


(a)


(b)

FIGURE 8-8 Iron, Fe, and oxygen, O_2, combine to form two different oxides: (a) iron(II) oxide, FeO, and (b) iron(III) oxide, Fe_2O_3.


(a)

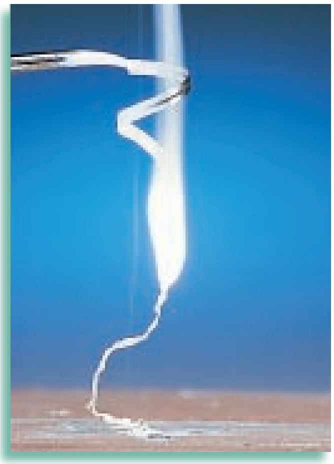

(b)

FIGURE 8-7 Magnesium, Mg, pictured in (a), undergoes a synthesis reaction with oxygen, O_2, in the air to produce magnesium oxide, MgO, as shown in (b).

Reactions of Metals with Halogens

Most metals react with the Group 17 elements, the halogens, to form either ionic or covalent compounds. For example, Group 1 metals react with halogens to form ionic compounds with the formula MX, where M is the metal and X is the halogen. Examples of this type of synthesis reaction include the reactions of sodium with chlorine and potassium with iodine.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$
$$2K(s) + I_2(g) \longrightarrow 2KI(s)$$

Group 2 metals react with the halogens to form ionic compounds with the formula MX_2.

$$Mg(s) + F_2(g) \longrightarrow MgF_2(s)$$
$$Sr(s) + Br_2(l) \longrightarrow SrBr_2(s)$$

The halogens undergo synthesis reactions with many different metals. Fluorine in particular is so reactive that it combines with almost all metals. For example, fluorine reacts with sodium to produce sodium fluoride. Similarly, it reacts with cobalt to form cobalt(III) fluoride and with uranium to form uranium(VI) fluoride.

$$2Na(s) + F_2(g) \longrightarrow 2NaF(s)$$
$$2Co(s) + 3F_2(g) \longrightarrow 2CoF_3(s)$$
$$U(s) + 3F_2(g) \longrightarrow UF_6(g)$$

Sodium fluoride, NaF, is added to municipal water supplies in trace amounts to provide fluoride ions, which help to prevent tooth decay in the people who drink the water. Cobalt(III) fluoride, CoF_3, is a strong fluorinating agent. And natural uranium is converted to uranium(VI) fluoride, UF_6, as the first step in the production of uranium for use in nuclear power plants.

FIGURE 8-9 Calcium hydroxide, a base, can be used to *neutralize* hydrochloric acid in your stomach. You will read more about acids, bases, and neutralization in Chapter 15.

Synthesis Reactions with Oxides

Active metals are highly reactive metals. Oxides of active metals react with water to produce metal hydroxides. For example, calcium oxide reacts with water to form calcium hydroxide, an ingredient in some stomach antacids.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(s)$$

Calcium oxide, CaO, also known as lime or quicklime, is manufactured in large quantities. The addition of water to lime to produce $Ca(OH)_2$, which is also known as slaked lime, is a crucial step in the setting of cement.

Many oxides of nonmetals in the upper right portion of the periodic table react with water to produce oxyacids. For example, sulfur dioxide, SO_2, reacts with water to produce sulfurous acid.

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

In air polluted with SO_2, sulfurous acid further reacts with oxygen to form sulfuric acid, one of the main ingredients in *acid rain*.

$$2H_2SO_3(aq) + O_2(g) \longrightarrow 2H_2SO_4(aq)$$

Certain metal oxides and nonmetal oxides react with each other in synthesis reactions to form salts. For example, calcium sulfite is formed by the reaction of calcium oxide and sulfur dioxide.

$$CaO(s) + SO_2(g) \longrightarrow CaSO_3(s)$$

Decomposition Reactions

In a **decomposition reaction,** *a single compound undergoes a reaction that produces two or more simpler substances.* Decomposition reactions are the opposite of synthesis reactions and are represented by the following general equation.

$$AX \longrightarrow A + X$$

AX is a compound. A and X can be elements or compounds.

Most decomposition reactions take place only when energy in the form of electricity or heat is added. Examples of several types of decomposition reactions are given in the following sections.

Decomposition of Binary Compounds

The simplest kind of decomposition reaction is the decomposition of a binary compound into its elements. We have already examined one example of a decomposition reaction. Figure 8-5 on page 250 shows that passing an electric current through water will decompose the water into its constituent elements, hydrogen and oxygen.

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

The decomposition of a substance by an electric current is called **electrolysis.**

Oxides of the less-active metals, which are located in the lower center of the periodic table, decompose into their elements when heated. Joseph Priestley discovered oxygen through such a decomposition reaction in 1774, when he heated mercury(II) oxide to produce mercury and oxygen.

$$2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$$

This reaction is shown in Figure 8-10 on page 260.

FIGURE 8-10 When mercury(II) oxide (the red-orange substance in the bottom of the test tube) is heated, it decomposes into oxygen and metallic mercury, which can be seen as droplets on the inside wall of the test tube.

Decomposition of Metal Carbonates

When a metal carbonate is heated, it breaks down to produce a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes to produce calcium oxide and carbon dioxide.

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

Decomposition of Metal Hydroxides

All metal hydroxides except those containing Group 1 metals decompose when heated to yield metal oxides and water. For example, calcium hydroxide decomposes to produce calcium oxide and water.

$$Ca(OH)_2(s) \xrightarrow{\Delta} CaO(s) + H_2O(g)$$

Decomposition of Metal Chlorates

When a metal chlorate is heated, it decomposes to produce a metal chloride and oxygen. For example, potassium chlorate, $KClO_3$, decomposes in the presence of the catalyst $MnO_2(s)$ to produce potassium chloride and oxygen.

$$2KClO_3(s) \xrightarrow[MnO_2(s)]{\Delta} 2KCl(s) + 3O_2(g)$$

Decomposition of Acids

Certain acids decompose into nonmetal oxides and water. Carbonic acid is unstable and decomposes readily at room temperature to produce carbon dioxide and water.

$$H_2CO_3(aq) \longrightarrow CO_2(g) + H_2O(l)$$

When heated, sulfuric acid decomposes into sulfur trioxide and water.

$$H_2SO_4(aq) \xrightarrow{\Delta} SO_3(g) + H_2O(l)$$

Sulfurous acid, H_2SO_3, decomposes similarly.

Single-Replacement Reactions

In a **single-replacement reaction,** *also known as a* **displacement reaction,** *one element replaces a similar element in a compound.* Many single-replacement reactions take place in aqueous solution. The amount of energy involved in this type of reaction is usually smaller than the amount involved in synthesis or decomposition reactions. Single-replacement reactions can be represented by the following general equations.

$$A + BX \longrightarrow AX + B$$
or
$$Y + BX \longrightarrow BY + X$$

A, B, X, and Y are elements. AX, BX, and BY are compounds.

Replacement of a Metal in a Compound by Another Metal

Aluminum is more active than lead. When solid aluminum is placed in aqueous lead(II) nitrate, $Pb(NO_3)_2(aq)$, the aluminum replaces the lead. Solid lead and aqueous aluminum nitrate are formed.

$$2Al(s) + 3Pb(NO_3)_2(aq) \longrightarrow 3Pb(s) + 2Al(NO_3)_3(aq)$$

Replacement of Hydrogen in Water by a Metal

The most-active metals, such as those in Group 1, react vigorously with water to produce metal hydroxides and hydrogen. For example, sodium reacts with water to form sodium hydroxide and hydrogen gas.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

Less-active metals, such as iron, react with steam to form a metal oxide and hydrogen gas.

$$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$$

Replacement of Hydrogen in an Acid by a Metal

The more-active metals react with certain acidic solutions, such as hydrochloric acid and dilute sulfuric acid, replacing the hydrogen in the acid. The reaction products are a metal compound (a salt) and hydrogen gas. For example, when solid magnesium reacts with hydrochloric acid, as shown in Figure 8-11, the reaction products are hydrogen gas and aqueous magnesium chloride.

$$Mg(s) + 2HCl(aq) \longrightarrow H_2(g) + MgCl_2(aq)$$

FIGURE 8-11 In this single-replacement reaction, the hydrogen in hydrochloric acid, HCl, is replaced by magnesium, Mg.

Replacement of Halogens

In another type of single-replacement reaction, one halogen replaces another halogen in a compound. Fluorine is the most-active halogen. As such, it can replace any of the other halogens in their compounds. Each halogen is less active than the one above it in the periodic table. Therefore, in Group 17 each element can replace any element below it, but not any element above it. For example, while chlorine can replace bromine in potassium bromide, it cannot replace fluorine in potassium fluoride. The reaction of chlorine with potassium bromide produces bromine and potassium chloride, whereas the combination of fluorine and sodium chloride produces sodium fluoride and solid chlorine.

$$Cl_2(g) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(l)$$
$$F_2(g) + 2NaCl(aq) \longrightarrow 2NaF(aq) + Cl_2(g)$$
$$Br_2(l) + KCl(aq) \longrightarrow \text{no reaction}$$

internet**connect**

SCILINKS NSTA

TOPIC: Precipitate
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2083

Double-Replacement Reactions

In **double-replacement reactions,** *the ions of two compounds exchange places in an aqueous solution to form two new compounds.* One of the compounds formed is usually a precipitate, an insoluble gas that bubbles out of the solution, or a molecular compound, usually water. The other compound is often soluble and remains dissolved in solution. A double-replacement reaction is represented by the following general equation.

$$AX + BY \longrightarrow AY + BX$$

A, X, B, and Y in the reactants represent ions. AY and BX represent ionic or molecular compounds.

Formation of a Precipitate

The formation of a precipitate occurs when the cations of one reactant combine with the anions of another reactant to form an insoluble or slightly soluble compound. For example, when an aqueous solution of potassium iodide is added to an aqueous solution of lead(II) nitrate, the yellow precipitate lead(II) iodide forms. This is shown in Figure 8-12.

$$2KI(aq) + Pb(NO_3)_2(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$$

The precipitate forms as a result of the very strong attractive forces between the Pb^{2+} cations and the I^- anions. The other product is the water-soluble salt potassium nitrate, KNO_3. The potassium and nitrate ions do not take part in the reaction. They remain in solution as aqueous ions. The guidelines that help identify which ions form a precipitate and which ions remain in solution are developed in Chapter 14.

FIGURE 8-12 The double-replacement reaction between aqueous lead(II) nitrate, $Pb(NO_3)_2(aq)$, and aqueous potassium iodide, $KI(aq)$, yields the precipitate lead(II) iodide, $PbI_2(s)$.

Formation of a Gas

In some double-replacement reactions, one of the products is an insoluble gas that bubbles out of the mixture. For example, iron(II) sulfide reacts with hydrochloric acid to form hydrogen sulfide gas and iron(II) chloride.

$$FeS(s) + 2HCl(aq) \longrightarrow H_2S(g) + FeCl_2(aq)$$

Formation of Water

In some double-replacement reactions, a very stable molecular compound, such as water, is one of the products. For example, hydrochloric acid reacts with an aqueous solution of sodium hydroxide to yield aqueous sodium chloride and water.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

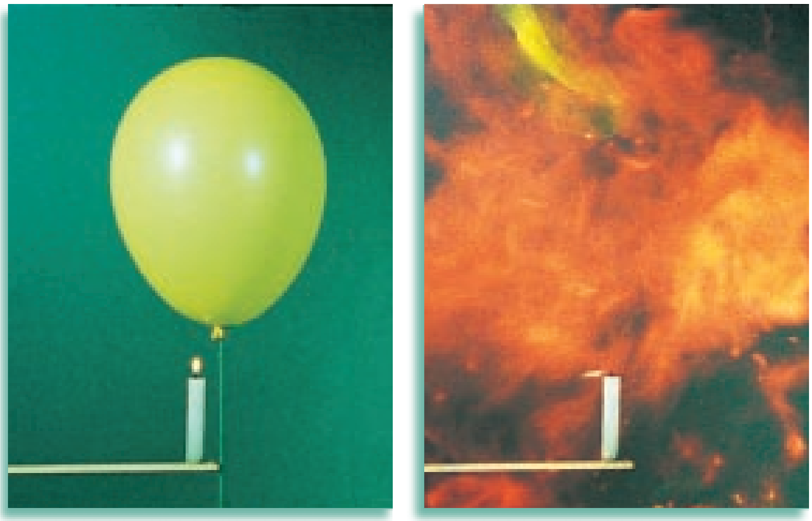
Combustion Reactions

In a **combustion reaction,** *a substance combines with oxygen, releasing a large amount of energy in the form of light and heat.* The combustion of hydrogen is shown below in Figure 8-13. The reaction's product is water vapor.

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

The burning of natural gas, propane, gasoline, and wood are also examples of combustion reactions. For example, the burning of propane, C_3H_8, results in the production of carbon dioxide and water vapor.

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



(a) (b)

FIGURE 8-13 (a) The candle supplies heat to the hydrogen and oxygen in the balloon, triggering the explosive combustion reaction shown in (b).

Quick LAB

Balancing Equations Using Models

Wear Safety Goggles and an Apron.

Materials

- large and small gumdrops in at least four different colors
- toothpicks

Question

How can molecular models and formula-unit ionic models be used to balance chemical equations and classify chemical reactions?

Procedure

Examine the partial equations in Groups A–E. Using different-colored gumdrops to represent atoms of different elements, make models of the reactions by connecting the appropriate "atoms" with toothpicks. Use your models to (1) balance equations (a) and (b) in each group, (2) determine the products for reaction (c) in each group, and (3) complete and balance each equation (c). Finally, (4) classify each group of reactions by type.

Group A

a. $H_2 + Cl_2 \longrightarrow HCl$

b. $Mg + O_2 \longrightarrow MgO$

c. $BaO + H_2O \longrightarrow$ _____

Group B

a. $H_2CO_3 \longrightarrow CO_2 + H_2O$

b. $KClO_3 \longrightarrow KCl + O_2$

c. $H_2O \xrightarrow{\text{electricity}}$ _____

Group C

a. $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$

b. $KI + Br_2 \longrightarrow KBr + I_2$

c. $Zn + HCl \longrightarrow$ _____

Group D

a. $AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$

b. $FeS + HCl \longrightarrow FeCl_2 + H_2S$

c. $H_2SO_4 + KOH \longrightarrow$ _____

Group E

a. $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

b. $CO + O_2 \longrightarrow CO_2$

c. $C_3H_8 + O_2 \longrightarrow$ _____

SECTION REVIEW

1. List five types of chemical reactions.
2. Classify each of the following reactions as synthesis, decomposition, single-replacement, double-replacement, or combustion:
   a. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
   b. $2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$
   c. $2NaNO_3(s) \longrightarrow 2NaNO_2(s) + O_2(g)$
   d. $2C_6H_{14}(l) + 19O_2(g) \longrightarrow 12CO_2(g) + 14H_2O(l)$
   e. $NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$
   f. $BaO(s) + H_2O(l) \longrightarrow Ba(OH)_2(aq)$
   g. $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$
3. For each of the following reactions, identify the missing reactant(s) or products(s) and then balance the resulting equation. Note that each empty slot may require one or more substances.
   a. synthesis: _____ $\longrightarrow Li_2O$
   b. decomposition: $Mg(ClO_3)_2 \longrightarrow$ _____
   c. single-replacement: $Na + H_2O \longrightarrow$ _____
   d. double-replacement: $HNO_3 + Ca(OH)_2 \longrightarrow$ _____
   e. combustion: $C_5H_{12} + O_2 \longrightarrow$ _____
4. For each of the following reactions, write the missing product(s) and then balance the resulting equation. Identify each reaction by type.
   a. $Br_2 + KI \longrightarrow$ _____
   b. $Zn + HCl \longrightarrow$ _____
   c. $Ca + Cl_2 \longrightarrow$ _____
   d. $NaClO_3 \xrightarrow{\Delta}$ _____
   e. $C_7H_{14} + O_2 \longrightarrow$ _____
   f. $CuCl_2 + Na_2S \longrightarrow$ _____

SECTION 8-3

Activity Series of the Elements

OBJECTIVES

- Explain the significance of an activity series.
- Use an activity series to predict whether a given reaction will occur and what the products will be.

The ability of an element to react is referred to as the element's *activity*. The more readily an element reacts with other substances, the greater its activity is. *An* **activity series** *is a list of elements organized according to the ease with which the elements undergo certain chemical reactions*. For metals, greater activity means a greater ease of *loss* of electrons, to form positive ions. For nonmetals, greater activity means a greater ease of *gain* of electrons, to form negative ions.

The order in which the elements are listed is usually determined by single-replacement reactions. The most-active element, placed at the top in the series, can replace each of the elements below it from a compound in a single-replacement reaction. An element farther down can replace any element below it but not any above it. For example, in the discussion of single-replacement reactions in Section 8-2, it was noted that each halogen will react to replace any halogen listed below it in the periodic table. Therefore, an activity series for the Group 17 elements lists them in the same order, from top to bottom, as they appear in the periodic table. This is shown in Table 8-3 on page 266.

As mentioned in Section 8-1, the fact that a chemical equation can be written does not necessarily mean that the reaction it represents will actually take place. Activity series are used to help predict whether certain chemical reactions will occur. For example, according to the activity series for metals in Table 8-3, aluminum replaces zinc. Therefore, we could predict that the following reaction does occur.

$$2Al(s) + 3ZnCl_2(aq) \longrightarrow 3Zn(s) + 2AlCl_3(aq)$$

Cobalt, however, cannot replace sodium. Therefore, we write the following.

$$Co(s) + 2NaCl(aq) \longrightarrow \text{no reaction}$$

It is important to remember that like many other aids used to predict the products of chemical reactions, activity series are based on experiment. The information that they contain is used as a general guide for predicting reaction outcomes. For example, the activity series reflects the fact that some metals (potassium, for example) react vigorously with water and acids, replacing hydrogen to form new compounds. Other metals, such as iron or zinc, replace hydrogen in acids such as hydrochloric acid but react with water only when the water is hot

internet**connect**

SCILINKS NSTA

TOPIC: Activity series
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2084

enough to become steam. Nickel, on the other hand, will replace hydrogen in acids but will not react with steam at all. And gold will not react with acid or water, either as a liquid or as steam. Such experimental observations are the basis for the activity series shown in Table 8-3.

TABLE 8-3 Activity Series of the Elements

Activity of metals		Activity of halogen nonmetals
Li Rb K Ba Sr Ca Na	React with cold H_2O and acids, replacing hydrogen. React with oxygen, forming oxides.	F_2 Cl_2 Br_2 I_2
Mg Al Mn Zn Cr Fe Cd	React with steam (but not cold water) and acids, replacing hydrogen. React with oxygen, forming oxides.	
Co Ni Sn Pb	Do not react with water. React with acids, replacing hydrogen. React with oxygen, forming oxides.	
H_2 Sb Bi Cu Hg	React with oxygen, forming oxides.	
Ag Pt Au	Fairly unreactive, forming oxides only indirectly.	

SAMPLE PROBLEM 8-6

Using the activity series shown in Table 8-3, explain whether each of the possible reactions listed below will occur. For those reactions that will occur, predict what the products will be.

a. $\mathbf{Zn(}s\mathbf{) + H_2O(}l\mathbf{) \xrightarrow{50°C}}$ _____
b. $\mathbf{Sn(}s\mathbf{) + O_2(}g\mathbf{) \longrightarrow}$ _____
c. $\mathbf{Cd(}s\mathbf{) + Pb(NO_3)_2(}aq\mathbf{) \longrightarrow}$ _____
d. $\mathbf{Cu(}s\mathbf{) + HCl(}aq\mathbf{) \longrightarrow}$ _____

SOLUTION

a. This is a reaction between a metal and water at 50°C. Zinc reacts with water only when it is hot enough to be steam. Therefore, no reaction will occur.

b. Any metal more active than silver will react with oxygen to form an oxide. Tin is above silver in the activity series. Therefore, a reaction will occur, and the product will be a tin oxide, either SnO or SnO_2.

c. An element will replace any element below it in the activity series from a compound in aqueous solution. Cadmium is above lead, and therefore a reaction will occur to produce lead, Pb, and cadmium nitrate, $Cd(NO_3)_2$.

d. Any metal more active than hydrogen will replace hydrogen from an acid. Copper is not above hydrogen in the series. Therefore, no reaction will occur.

PRACTICE

1. Using the activity series shown in Table 8-3, predict whether each of the possible reactions listed below will occur. For the reactions that will occur, write the products and balance the equation.

a. $Cr(s) + H_2O(l) \longrightarrow$ _____

b. $Pt(s) + O_2(g) \longrightarrow$ _____

c. $Cd(s) + 2HBr(aq) \longrightarrow$ _____

d. $Mg(s) + \text{steam} \longrightarrow$ _____

Answer

1. a. no

b. no

c. yes; $2Cd(s) + 2HBr(aq) \longrightarrow 2CdBr(aq) + H_2(g)$

d. yes; $Mg(s) + 2H_2O(g) \longrightarrow Mg(OH)_2(aq) + H_2(g)$

2. Identify the element that replaces hydrogen from acids but cannot replace tin from its compounds.

2. Pb

3. According to Table 8-3, what is the most-active transition metal?

3. Mn

SECTION REVIEW

1. How is the activity series useful in predicting chemical behavior?

2. Based on the activity series, predict whether each of the following possible reactions listed will occur:

a. $Ni(s) + H_2O(l) \longrightarrow$ _____

b. $Br_2(l) + KI(aq) \longrightarrow$ _____

c. $Au(s) + HCl(aq) \longrightarrow$ _____

d. $Cd(s) + HCl(aq) \longrightarrow$ _____

e. $Mg(s) + Co(NO_3)_2(aq) \longrightarrow$ _____

3. For each of the reactions in item 2 that will occur, write the products and balance the equation.

Acid Water—A Hidden Menace

When purchasing a home with its own well, it is common practice to have the water in the well tested. Usually, the purpose of the tests is to indicate the presence of disease-causing microorganisms. Rarely is the water's acidity measured.

Many people are unaware of their water's pH value (see Chapter 16) until they are confronted with such phenomena as a blue ring materializing around a porcelain sink drain, a water heater suddenly giving out, or tropical fish that keep dying. Each of these events could be traced to acidic water, which can also be a cause of lead poisoning.

The possibility of lead poisoning from home water supplies has gone largely unreported. Many older homes still have lead pipes in their plumbing, while most modern homes use copper piping. All pipe joints, however, are sealed with lead solder. Highly acidic water can leach out both the lead from the solder joints and copper from the pipes themselves, which turns the sink drain blue. In addition, people who are in the habit of filling their kettles in the morning without letting the tap run awhile first could be adding a number of unwanted chemicals to their tea or coffee.

Lead poisoning is of particular concern in young children. The absorption rate of lead in the intestinal tract of a child is much higher than that of an adult, and lead poisoning can permanently impair a child's rapidly growing nervous system. The good news is that lead poisoning and other effects of acidic water in the home can be easily prevented. Here's what you can do about it:

1. Monitor the pH of your water on a regular basis, especially if you have well water. This can easily be done with pH test kits (see photograph) that are sold in hardware or pet stores—many tropical fish are intolerant of water with a pH that is either too high (basic) or too low (acidic). The pH of most municipal water supplies should already be regulated, but it doesn't hurt to check.
2. In the morning, let your water tap run for about half a minute before you fill your kettle or drink the water. If the water is acidic, the first flush of water will have the highest concentration of lead and copper ions.
3. Installing an alkali-injection pump is a low-cost, low-maintenance solution that can save your plumbing and lessen the risk of lead poisoning from your own water supply. The pump injects a small amount of an alkali (usually potassium carbonate or sodium carbonate) in your water-holding tank each time you activate your well's pump. This effectively neutralizes the acidity of your water. The reaction below shows the neutralizing effect of potassium carbonate on well water that has been made acidic by acid rain.

$$K_2CO_3(aq) + H_2SO_4(aq) \longrightarrow K_2SO_4(aq) + CO_2(g) + H_2O(l)$$

The pH of your home's water supply can be easily monitored using a test kit like the one shown here.

internet connect

SCILINKS NSTA

TOPIC: Acid water
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2085

CHAPTER 8 REVIEW

CHAPTER SUMMARY

8-1

- Five observations that suggest a chemical reaction is taking place are the evolution of heat or light, the production of gas, a change in color, and the formation of a precipitate.
- A balanced chemical equation represents, with symbols and formulas, the identities and relative amounts of reactants and products in a chemical reaction.

Vocabulary

chemical equation (241)
coefficient (243)
formula equation (244)
precipitate (242)
reversible reaction (246)
word equation (243)

8-2

- Synthesis reactions are represented by the general equation $A + X \longrightarrow AX$.
- Decomposition reactions are represented by the general equation $AX \longrightarrow A + X$.
- Single-replacement reactions are represented by the general equations $A + BX \longrightarrow AX + B$ and $Y + BX \longrightarrow BY + X$.
- Double-replacement reactions are represented by the general equation $AX + BY \longrightarrow AY + BX$.

Vocabulary

combustion reaction (263)
composition reaction (256)
decomposition reaction (259)
displacement reaction (261)
double-replacement reaction (262)
electrolysis (259)
single-replacement reaction (261)
synthesis reaction (256)

8-3

- Activity series list the elements in order of their chemical reactivity and are useful in predicting whether a chemical reaction will occur.
- Chemists determine activity series through experiments.

Vocabulary

activity series (265)

REVIEWING CONCEPTS

1. List four observations that indicate that a chemical reaction may be taking place. (8-1)

2. List the three requirements for a correctly written chemical equation. (8-1)

3. a. What is meant by the term *coefficient* in relation to a chemical equation?
b. How does the presence of a coefficient affect the number of atoms of each type in the formula that it precedes? (8-1)

4. Give an example of a word equation, a formula equation, and a chemical equation. (8-1)

5. What quantitative information is revealed by a chemical equation? (8-1)

6. What limitations are associated with the use of both word and formula equations? (8-1)

7. Define each of the following:
a. aqueous solution
b. catalyst
c. reversible reaction (8-1)

8. Write formulas for each of the following compounds:
a. potassium hydroxide
b. calcium nitrate
c. sodium carbonate
d. carbon tetrachloride
e. magnesium bromide
f. sulfur dioxide
g. ammonium sulfate (8-1)

9. What four guidelines are useful in balancing an equation? (8-1)

10. How many atoms of each type are represented in each of the following?
   a. $3N_2$
   b. $2H_2O$
   c. $4HNO_3$
   d. $2Ca(OH)_2$
   e. $3Ba(ClO_3)_2$
   f. $5Fe(NO_3)_2$
   g. $4Mg_3(PO_4)_2$
   h. $2(NH_4)_2SO_4$
   i. $6Al_2(SeO_4)_3$
   j. $4C_3H_8$ (8-1)

11. Define and give general equations for the five basic types of chemical reactions introduced in Chapter 8. (8-2)

12. How are most decomposition reactions initiated? (8-2)

13. What is electrolysis? (8-2)

14. a. In what environment do many single-replacement reactions commonly occur?
   b. In general, how do single-replacement reactions compare with synthesis and decomposition reactions in terms of the amount of energy involved? (8-2)

15. a. What is meant by the *activity* of an element?
   b. How does this description differ for metals and nonmetals? (8-3)

16. a. What is an activity series of elements?
   b. What is the basis for the ordering of the elements in the activity series? (8-3)

17. a. What is the chemical principle upon which the activity series of metals is based?
   b. What is the significance of the distance between two metals in the activity series? (8-3)

PROBLEMS

Chemical Equations

18. Write the chemical equation that relates to each of the following word equations. Include symbols for physical states in the equation. (Hint: See Sample Problem 8-1.)
   a. solid zinc sulfide + oxygen gas $\longrightarrow$ solid zinc oxide + sulfur dioxide gas
   b. hydrochloric acid + aqueous magnesium hydroxide $\longrightarrow$ aqueous magnesium chloride + water
   c. nitric acid + aqueous calcium hydroxide $\longrightarrow$ aqueous calcium nitrate + water

19. Translate each of the following chemical equations into a sentence. (Hint: See Sample Problem 8-2.)
   a. $2ZnS(s) + 3O_2(g) \longrightarrow 2ZnO(s) + 2SO_2(g)$
   b. $CaH_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + 2H_2(g)$
   c. $AgNO_3(aq) + KI(aq) \longrightarrow AgI(s) + KNO_3(aq)$

20. Balance each of the following:
   a. $H_2 + Cl_2 \longrightarrow HCl$
   b. $Al + Fe_2O_3 \longrightarrow Al_2O_3 + Fe$
   c. $Pb(CH_3COO)_2 + H_2S \longrightarrow PbS + CH_3COOH$

21. The following equations are incorrect in some way. Identify and correct each error, and then balance each equation.
   a. $Li + O_2 \longrightarrow LiO_2$
   b. $H_2 + Cl_2 \longrightarrow H_2Cl_2$
   c. $MgCO_3 \longrightarrow MgO_2 + CO_2$
   d. $NaI + Cl_2 \longrightarrow NaCl + I$

22. Write chemical equations for each of the following sentences:
   a. Aluminum reacts with oxygen to produce aluminum oxide.
   b. Phosphoric acid, H_3PO_4, is produced through the reaction between tetraphosphorus decoxide and water.
   c. Iron(III) oxide reacts with carbon monoxide to produce iron and carbon dioxide.

23. Carbon tetrachloride is used as an intermediate chemical in the manufacture of other chemicals. It is prepared in liquid form by reacting chlorine gas with methane gas. Hydrogen chloride gas is also formed in this reaction. Write the balanced chemical equation for the production of carbon tetrachloride. (Hint: See Sample Problems 8-3 and 8-4.)

24. For each of the following synthesis reactions, identify the missing reactant(s) or product(s), and then balance the resulting equation:
   a. $Mg + \_\_\_\_ \longrightarrow MgO$
   b. $\_\_\_\_ + O_2 \longrightarrow Fe_2O_3$
   c. $Li + Cl_2 \longrightarrow \_\_\_\_$
   d. $Ca + \_\_\_\_ \longrightarrow CaI_2$

Types of Chemical Reactions

25. Complete the following synthesis reactions by writing both word and chemical equations for each:
a. sodium + oxygen $\longrightarrow$ _____
b. magnesium + fluorine $\longrightarrow$ _____

26. Complete and balance the equation for each of the following decomposition reactions:
a. $HgO \xrightarrow{\Delta}$
b. $H_2O(l) \xrightarrow{\text{electricity}}$
c. $Ag_2O \xrightarrow{\Delta}$
d. $CuCl_2 \xrightarrow{\text{electricity}}$

27. Complete and balance the equations for each of the following single-replacement reactions:
a. $Zn + Pb(NO_3)_2 \longrightarrow$ _____
b. $Al + Hg(CH_3COO)_2 \longrightarrow$ _____
c. $Al + NiSO_4 \longrightarrow$ _____
d. $Na + H_2O \longrightarrow$ _____

28. Complete and balance the equations for the following double-replacement reactions:
a. $AgNO_3(aq) + NaCl(aq) \longrightarrow$ _____
b. $Mg(NO_3)_2(aq) + KOH(aq) \longrightarrow$ _____
c. $LiOH(aq) + Fe(NO_3)_3(aq) \longrightarrow$ _____

29. Complete and balance the equation for each of the following combustion reactions:
a. $CH_4 + O_2 \longrightarrow$ _____
b. $C_3H_6 + O_2 \longrightarrow$ _____
c. $C_5H_{12} + O_2 \longrightarrow$ _____

30. Write and balance each of the following equations, and then identify each by type:
a. hydrogen + iodine $\longrightarrow$ hydrogen iodide
b. lithium + hydrochloric acid $\longrightarrow$ lithium chloride + hydrogen
c. sodium carbonate $\longrightarrow$ sodium oxide + carbon dioxide
d. mercury(II) oxide $\longrightarrow$ mercury + oxygen
e. magnesium hydroxide $\longrightarrow$ magnesium oxide + water

31. Identify the compound that could undergo decomposition to produce the following products, and then balance the final equation:
a. magnesium oxide and water
b. lead(II) oxide and water
c. lithium chloride and oxygen
d. barium chloride and oxygen
e. nickel chloride and oxygen

32. In each of the following combustion reactions, identify the missing reactant(s), product(s), or both, and then balance the resulting equation:
a. $C_3H_8 +$ _____ $\longrightarrow$ _____ $+ H_2O$
b. _____ $+ 8O_2 \longrightarrow 5CO_2 + 6H_2O$
c. $C_2H_5OH +$ _____ $\longrightarrow$ _____ + _____

33. Complete and balance each of the following reactions observed to occur, and then identify each by type:
a. zinc + sulfur $\longrightarrow$ _____
b. calcium + sodium nitrate $\longrightarrow$ _____
c. silver nitrate + potassium iodide $\longrightarrow$ _____
d. sodium iodide $\xrightarrow{\Delta}$ _____
e. toluene, C_7H_8 + oxygen $\longrightarrow$ _____
f. nonane, C_9H_{20} + oxygen $\longrightarrow$ _____

Activity Series

34. Based on the activity series of metals and halogens, which element within each pair is more likely to replace the other in a compound?
a. K and Na
b. Al and Ni
c. Bi and Cr
d. Cl and F
e. Au and Ag
f. Cl and I
g. Fe and Sr
h. I and F

35. Using the activity series in Table 8-3 on page 266, predict whether each of the possible reactions listed below will occur. For the reactions that will occur, write the products and balance the equation.
a. $Ni(s) + CuCl_2(aq) \longrightarrow$ _____
b. $Zn(s) + Pb(NO_3)_2(aq) \longrightarrow$ _____
c. $Cl_2(g) + KI(aq) \longrightarrow$ _____
d. $Cu(s) + FeSO_4(aq) \longrightarrow$ _____
e. $Ba(s) + H_2O(l) \longrightarrow$ _____

36. Use the activity series to predict whether each of the following synthesis reactions will occur, and write the chemical equations for those predicted to occur:
a. $Ca(s) + O_2(g) \longrightarrow$ _____
b. $Ni(s) + O_2(g) \longrightarrow$ _____
c. $Au(s) + O_2(g) \longrightarrow$ _____

MIXED REVIEW

37. Ammonia reacts with oxygen to yield nitrogen and water.

$$4NH_3(g) + 3O_2(g) \longrightarrow 2N_2(g) + 6H_2O(l)$$

Given this chemical equation, as well as the number of moles of the reactant or product indicated below, determine the number of moles of all remaining reactants and products.

a. 3.0 mol O_2
b. 8.0 mol NH_3
c. 1.0 mol N_2
d. 0.40 mol H_2O

38. Complete the following synthesis reactions by writing both the word and chemical equation for each:

a. potassium + chlorine $\longrightarrow$ _____
b. hydrogen + iodine $\longrightarrow$ _____
c. magnesium + oxygen $\longrightarrow$ _____

39. Use the activity series to predict which metal, Sn, Mn, or Pt, would be the best choice as a container for an acid.

40. Aqueous sodium hydroxide is produced commercially by the electrolysis of aqueous sodium chloride. Hydrogen and chlorine gases are also produced. Write the balanced chemical equation for the production of sodium hydroxide. Include the physical states of the reactants and products.

41. Balance each of the following:

a. $Ca(OH)_2 + (NH_4)_2SO_4 \longrightarrow CaSO_4 + NH_3 + H_2O$
b. $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$
c. $Cu_2S + O_2 \longrightarrow Cu_2O + SO_2$
d. $Al + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2$

42. Use the activity series to predict whether each of the following reactions will occur, and write the balanced chemical equations for those predicted to occur:

a. $Al(s) + O_2(g) \longrightarrow$ _____
b. $Pb(s) + ZnCl_2(s) \longrightarrow$ _____
c. $Rb(s) + Zn(NO_3)_2(aq) \longrightarrow$ _____

43. Complete and balance the equations for the following reactions, and identify the type of reaction each represents:

a. $(NH_4)_2S(aq) + ZnCl_2(aq) \longrightarrow$ _____ $+ ZnS(s)$
b. $Al(s) + Pb(NO_3)_2(aq) \longrightarrow$ _____
c. $Ba(s) + H_2O(l) \longrightarrow$ _____
d. $Cl_2(g) + KBr(aq) \longrightarrow$
e. $NH_3(g) + O_2(g) \xrightarrow{Pt} NO(g) + H_2O(l)$
f. $H_2O(l) \longrightarrow H_2(g) + O_2(g)$

44. Write and balance each of the following equations, and then identify each by type:

a. copper + chlorine $\longrightarrow$ copper(II) chloride
b. calcium chlorate $\longrightarrow$ calcium chloride + oxygen
c. lithium + water $\longrightarrow$ lithium hydroxide + hydrogen
d. lead(II) carbonate $\longrightarrow$ lead(II) oxide + carbon dioxide

45. How many moles of HCl can be made from 6.15 mol of H_2 and an excess of Cl_2?

46. What product is missing in the following equation?

$$MgO + 2HCl \longrightarrow MgCl_2 + \_\_\_\_\_$$

47. Balance the following equations:

a. $Pb(NO_3)_2(aq) + NaOH(aq) \longrightarrow Pb(OH)_2(s) + NaNO_3(aq)$
b. $C_{12}H_{22}O_{11}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$
c. $Al(OH)_3(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2O(l)$

48. Translate the following word equations into balanced chemical equations:

a. silver nitrate + potassium iodide $\longrightarrow$ silver iodide + potassium nitrate
b. nitrogen dioxide + water $\longrightarrow$ nitric acid + nitrogen monoxide
c. silicon tetrachloride + water $\longrightarrow$ silicon dioxide + hydrochloric acid

CRITICAL THINKING

49. **Inferring Relationships** Activity series are prepared by comparing single-displacement reactions between metals. Based on observations, the metals can be ranked by their ability to react. However, reactivity can be explained by the ease with which atoms of metals lose electrons. Using information from the activity series, identify the locations in the periodic table of the most-reactive metals and the least-reactive metals. Based on your knowledge of electron configurations and periodic trends, infer possible explanations for their reactivity and position in the periodic table.

50. Analyzing Results Formulate an activity series for the hypothetical elements A, J, Q, and Z using the reaction information provided.

$$A + ZX \longrightarrow AX + Z$$
$$J + ZX \longrightarrow \text{no reaction}$$
$$Q + AX \longrightarrow QX + A$$

HANDBOOK SEARCH

51. Find the common-reactions section for Group 1 metals in the *Elements Handbook*. Use this information to answer the following:
   a. Write a balanced chemical equation for the formation of rubidium hydroxide from rubidum oxide.
   b. Write a balanced chemical equation for the formation of cesium iodide.
   c. Classify the reactions you wrote in (a) and (b).
   d. Write word equations for the reactions you wrote in (a) and (b).

52. Find the common-reactions section for Group 13 in the *Elements Handbook*. Use this information to answer the following:
   a. Write a balanced chemical equation for the formation of gallium bromide prepared from hydrobromic acid.
   b. Write a balanced chemical equation for the formation of gallium oxide.
   c. Classify the reactions you wrote in (a) and (b).
   d. Write word equations for the reactions you wrote in (a) and (b).

53. Find the common-reactions section for Group 16 in the *Elements Handbook*. Use this information to answer the following:
   a. Write a balanced chemical equation for the formation of selenium trioxide.
   b. Write a balanced chemical equation for the formation of tellurium iodide.
   c. Classify the reactions you wrote in (a) and (b).
   d. Write word equations for the reactions you wrote in (a) and (b).

RESEARCH & WRITING

54. Trace the evolution of municipal water fluoridation. What advantages and disadvantages are associated with this practice?

55. Research how a soda-acid fire extinguisher works, and write the chemical equation for the reaction. Check your house and other structures for different types of fire extinguishers, and ask your local fire department to verify the effectiveness of each type of extinguisher.

ALTERNATIVE ASSESSMENT

56. Performance Assessment For one day, record situations that show evidence of a chemical change. Identify the reactants and the products, and determine whether there is proof of a chemical reaction. Classify each of the chemical reactions according to the common reaction types discussed in the chapter.

57. Design a set of experiments that will enable you to create an activity series for the elements composing the following metals and solutions:
   a. aluminum and aluminum chloride
   b. chromium and chromium(III) chloride
   c. iron and iron(II) chloride
   d. magnesium and magnesium chloride