CHAPTER 8

Chemical Equations and Reactions

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

Fireworks

Describing Chemical Reactions

SECTION 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 molecular or molar 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 1.**

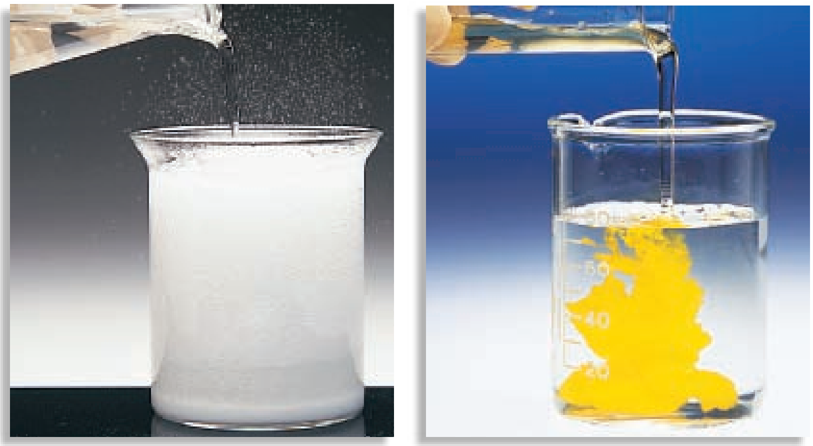
Indications of a Chemical Reaction

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 energy as 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 1** that the decomposition of ammonium dichromate is accompanied by the evolution of energy as 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 involve only heat or only light. But heat or light by itself is not necessarily a sign of chemical change, because many physical changes also involve either heat or light.

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

FIGURE 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 2a.**
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 2b.**
4. *Color change.* A change in color is often an indication of a chemical reaction.

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 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 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 balance numbers of atoms, add coefficients 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 261 indicates that 4 mol of water are produced for each mole of nitrogen and chromium(III) oxide that is produced.

extension

Chemistry in Action

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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.

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

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 3.**

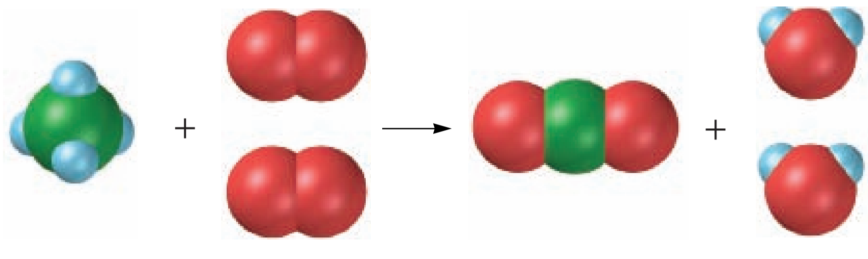
Additional Symbols Used in Chemical Equations

Table 2 on the next page 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,*

(a)

(b)



Molecules	$\underbrace{CH_4}$	+	$\underbrace{2O_2}$	$\longrightarrow$	CO_2	+	$2H_2O$
Atoms	$(\cancel{1C}, \cancel{4H})$	+	$(\cancel{4O})$	$\longrightarrow$	$(\cancel{1C}, \cancel{2O})$	+	$(\cancel{4H}, \cancel{2O})$

FIGURE 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 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.

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

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.

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

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 A

For more help, go to the Math Tutor at the end of this chapter.

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 B

Translate the following chemical equation into a sentence:

$$BaCl_2(aq) + Na_2CrO_4(aq) \longrightarrow BaCrO_4(s) + 2NaCl(aq)$$

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 barium chloride and sodium chromate react to produce a precipitate of barium chromate plus sodium chloride in aqueous solution.

PRACTICE *Answers in Appendix E*

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 1.**)
   c. Solid aluminum metal reacts with aqueous zinc chloride to produce solid zinc metal and aqueous aluminum chloride.
2. Translate the following chemical equations into sentences:
   a. $CS_2(l) + 3O_2(g) \longrightarrow CO_2(g) + 2SO_2(g)$
   b. $NaCl(aq) + AgNO_3(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$
3. Hydrazine, N_2H_4, is used as rocket fuel. Hydrazine reacts violently with oxygen to produce gaseous nitrogen and water. Write the balanced chemical equation.

extension

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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 4** in Chapter 7 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.

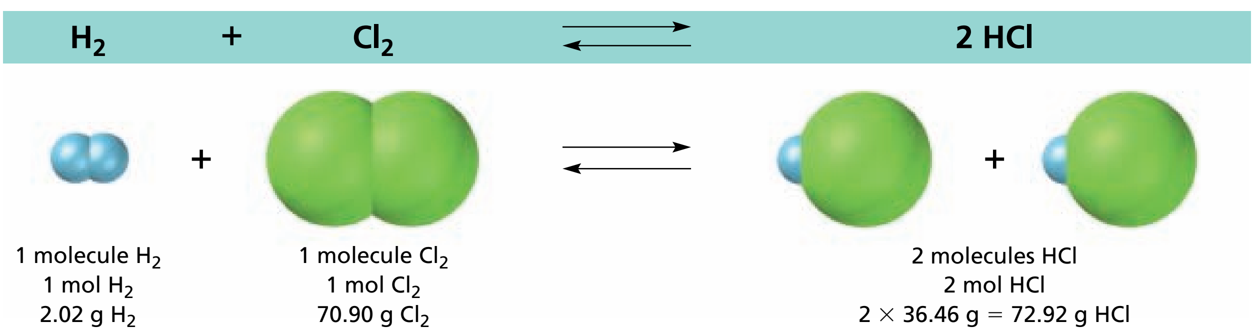
$$\cancel{1 \text{ mol } H_2} \times \frac{2.02 \text{ g } H_2}{\cancel{\text{mol } H_2}} = 2.02 \text{ g } H_2$$

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

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

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 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 the previous page 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 2 and 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 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 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 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.

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

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 C

*For more help, go to the **Math Tutor** at the end of this chapter.*

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 6. Write a balanced chemical equation for the reaction.

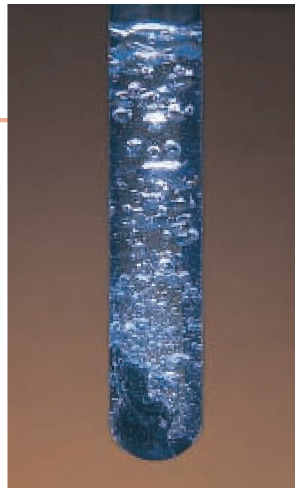


FIGURE 6 Solid zinc reacts with hydrochloric acid to form aqueous zinc chloride and hydrogen gas.

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 the previous page 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

Answers in Appendix E

1. Write word, formula, and balanced chemical equations for each of the following reactions:
   a. Solid magnesium and aqueous hydrochloric acid react to produce aqueous magnesium chloride and hydrogen gas.
   b. Aqueous nitric acid reacts with solid magnesium hydroxide to produce aqueous magnesium nitrate and water.
2. Solid calcium metal reacts with water to form aqueous calcium hydroxide and hydrogen gas. Write a balanced chemical equation for this reaction.

extension

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SAMPLE PROBLEM D

*For more help, go to the **Math Tutor** at the end of this chapter.*

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 E

*For more help, go to the **Math Tutor** at the end of this chapter.*

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.

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

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.

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

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

Answers in Appendix E

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.

extension

Go to **go.hrw.com** for more practice problems that ask you to write balanced chemical equations.

Keyword: HC6RXNX

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.

Critical Thinking

5. **INTEGRATING CONCEPTS** The reaction of vanadium(II) oxide with iron(III) oxide results in the formation of vanadium(V) oxide and iron(II) oxide. Write the balanced chemical equation.

Chemistry in Action

Carbon Monoxide Catalyst

Colorless, odorless, and deadly—carbon monoxide, "the silent killer," causes the deaths of hundreds of Americans every year. When fuel does not burn completely in a combustion process, carbon monoxide is produced. Often this occurs in a malfunctioning heater, furnace, or fireplace. When the carbon monoxide is inhaled, it bonds to the hemoglobin in the blood, leaving the body oxygen starved. Before people realize a combustion device is malfunctioning, it's often too late.

$$O_2Hb + CO \longrightarrow COHb + O_2$$

Carbon monoxide, CO, has almost 200 times the affinity to bind with the hemoglobin, Hb, in the blood as oxygen. This means that hemoglobin will bind to carbon monoxide rather than oxygen in the body. If enough carbon monoxide is present in the blood, it can be fatal.

Carbon monoxide poisoning can be prevented by installing filters that absorb the gas. After a time, however, filters become saturated, and then carbon monoxide can pass freely into the air. The best way to prevent carbon monoxide poisoning is not just to filter out the gas, but to eliminate it completely.

The solution came to research chemists at NASA who were working on a problem with a space-based laser. In order to operate properly, NASA's space-based carbon dioxide laser needed to be fed a continuous supply of CO_2. This was necessary because as a byproduct of its operation, the laser degraded some of the CO_2 into carbon monoxide and oxygen. To address this problem, NASA scientists developed a catalyst made of tin oxide and platinum that oxidized the waste carbon monoxide back into carbon dioxide. The NASA scientists then realized that this catalyst had the potential to be used in many applications here on Earth, including removing carbon monoxide from houses and other buildings.

Typically, a malfunctioning heater circulates the carbon monoxide it produces through its air intake system back into a dwelling space. Installing the catalyst in the air intake would oxidize any carbon monoxide to nontoxic carbon dioxide before it reentered the room.

"The form of our catalyst is a very thin coating on some sort of a support, or substrate as we call it," says NASA chemist David Schryer. "And that support, or substrate, can be any one of a number of things. The great thing about a catalyst is that the only thing that matters about it is its surface. So a catalyst can be incredibly thin and still be very effective."

The idea of using catalysts to oxidize gases is not a new one. Catalytic converters in cars oxidize carbon monoxide and unburned hydrocarbons to minimize pollution. Many substances are oxidized into new materials for manufacturing purposes. But both of these types of catalytic reactions occur at very high temperatures. NASA's catalyst is special, because it's able to eliminate carbon monoxide at room temperature.

According to David Schryer, low-temperature catalysts constitute a whole new class of catalysts with abundant applications for the future.

Questions

1. How did NASA's research on the space-based carbon dioxide laser result in a benefit for consumers?
2. According to the chemical reaction, if there are 4.5 mol of oxygenated hemoglobin present in an excess of carbon monoxide, how many moles of hemoglobin would release oxygen and bind to carbon monoxide? Explain your answer.

SCILINKS.
www.scilinks.org
Topic: Carbon Monoxide
Code: HC60219

SECTION 2

Objectives

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

Types of Chemical Reactions

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 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-displacement, double-displacement, and combustion reactions. 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 7** on the next page, 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.**

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.

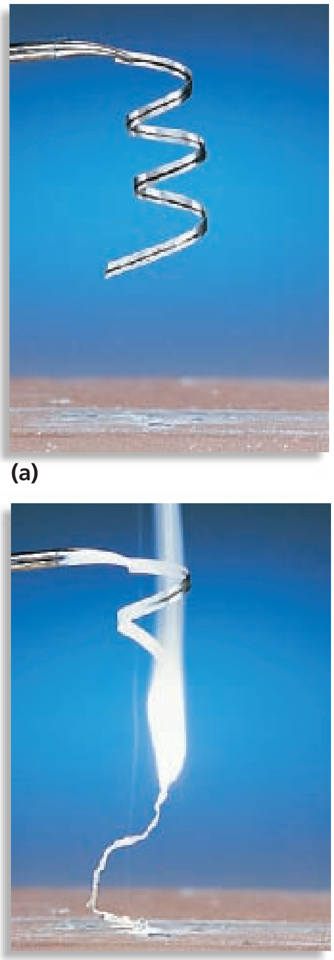
$$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$
$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

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 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).

(a) (b)

FIGURE 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.

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.

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

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

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

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.

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

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 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 14.

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, as shown in **Figure 9.**

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

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 5** on page 270 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 10** on the following page.

FIGURE 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-Displacement Reactions

In a **single-displacement reaction,** *also known as a replacement reaction, one element replaces a similar element in a compound.* Many single-displacement 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-displacement 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.

Displacement 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.

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

Displacement 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.

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

Displacement 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 11,** the reaction products are hydrogen gas and aqueous magnesium chloride.

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

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

Displacement of Halogens

In another type of single-displacement 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.

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

Double-Displacement Reactions

In **double-displacement 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-displacement 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 12.**

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

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

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 13.

Formation of a Gas

In some double-displacement 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-displacement 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.

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

> **Chemistry in Action**
>
> **Fluoridation and Tooth Decay**
>
> The main component of tooth enamel is a mineral called hydroxyapatite, $Ca_5(PO_4)_3OH$. Some foods contain acids or produce acids in the mouth, and acid dissolves tooth enamel, which leads to tooth decay. One way to help prevent tooth decay is by using fluoride. Fluoride reacts with hydroxyapatite in a double-displacement reaction. It displaces the OH^- group in hydroxyapatite to produce fluorapatite, $Ca_5(PO_4)_3F$. Studies show that calcium fluorapatite is about 20% less soluble than hydroxyapatite in acid. Therefore, fluoride lowers the incidence of tooth decay.

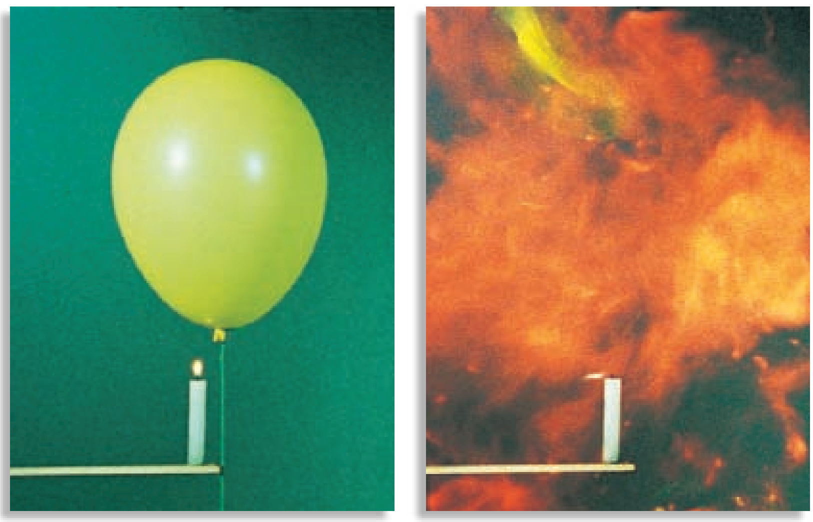
Combustion Reactions

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

QuickLAB

Wear safety goggles and an apron.

Balancing Equations Using Models

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. $C_3H_8 + O_2 \longrightarrow$ ______

Materials

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

SECTION REVIEW

1. List five types of chemical reactions.

2. Classify each of the following reactions as a synthesis, decomposition, single-displacement, double-displacement, or combustion reaction:

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. $2C_6H_{14}(l) + 19O_2(g) \longrightarrow 12CO_2(g) + 14H_2O(l)$

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. double displacement: $HNO_3 + Ca(OH)_2 \longrightarrow$ ________

d. 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. $NaClO_3 \xrightarrow{\Delta}$ ________

c. $C_7H_{14} + O_2 \longrightarrow$ ________

d. $CuCl_2 + Na_2S \longrightarrow$ ________

Critical Thinking

5. INFERRING RELATIONSHIPS In an experiment, an iron sample is oxidized to iron(III) oxide by oxygen, which is generated in the thermal decomposition of potassium chlorate. Write the two chemical reactions in the correct sequence.

Activity Series of the Elements

SECTION 3

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-displacement 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-displacement reaction. An element farther down can replace any element below it but not any above it. For example, in the discussion of single-displacement reactions in Section 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 3** on the next page.

As mentioned in Section 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 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

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Topic: Activity Series
Code: HC60019

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

TABLE 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 F

Using the activity series shown in Table 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. $Zn(s) + H_2O(l) \xrightarrow{50°C}$ _____
b. $Sn(s) + O_2(g) \longrightarrow$ _____
c. $Cd(s) + Pb(NO_3)_2(aq) \longrightarrow$ _____
d. $Cu(s) + HCl(aq) \longrightarrow$ _____

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

Answers in Appendix E

1. Using the activity series shown in **Table 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. $Mg(s)$ + steam $\longrightarrow$ ______

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

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

extension

Go to **go.hrw.com** for more practice problems that deal with activity series.

Keyword: HC6RXNX

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 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.

Critical Thinking

4. **PREDICTING OUTCOMES** A mixture contains cobalt metal, copper metal, and tin metal. This mixture is mixed with nickel nitrate. Which metals, if any, will react? Write the chemical equation for any reaction.

Chemistry in Action

Combustion Synthesis

What do aerospace materials, cutting tools, catalytic materials, ceramic engine parts, ball bearings, high-temperature superconductors, hydrogen storage, and fuel cells have in common? They are made of ceramics, composites, and other advanced materials.

Conventional techniques used to make these materials consist of a high-temperature furnace, with temperatures ranging from 500°C to 2000°C, to supply the energy needed for the reaction to take place. Because these furnaces may reach only 2000°C, it may take minutes to hours to convert reactants to solid-state products, and the mixtures are heated unevenly. As a result, flaws can be introduced into the structures, which can cause stress points in the materials.

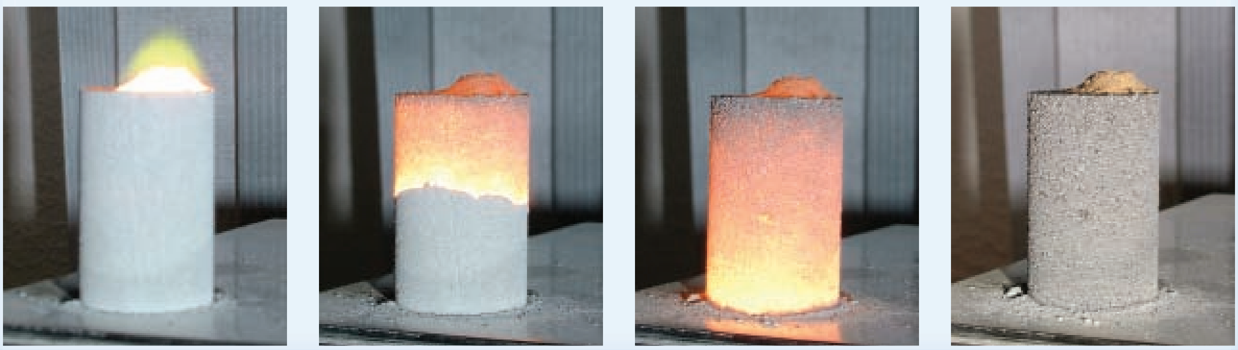
A different high-temperature technique is *combustion synthesis,* which generates its own energy to keep the reaction continuing. Once the reactant mixture is ignited, a heat wave moves through the sample, producing the solid-state product. The mixture can reach temperatures up to 4000°C, twice what is possible with conventional high-temperature furnaces. Combustion synthesis also allows reactions to be completed in just seconds. Hence, this technique produces the desired material faster and requires less supplied energy than conventional techniques do. In addition, the intense and quick heating produces materials that are chemically homogeneous. More than 500 compounds, such as lightweight and heat-resistant aerospace materials, are created by combustion synthesis.

In a typical combustion synthesis procedure, the reactant powders are mixed and then pressed into a cylindrical pellet. The pellet is ignited by an intense heat source, such as an electrically heated coil or a laser. Because the combustion-synthesis reaction is very exothermic, the reaction is self-propagating, and the process does not need any further input of energy. This type of self-propagation is called a *reaction wave,* in which the reaction propagates through the starting material in a self-sustained manner. Therefore, compared with conventional high-temperature methods, this technique is an energy-saving process. In addition, the high temperatures and short reaction times can produce materials that would not be synthesized under conventional conditions. Currently, scientists are studying reaction waves, including how they move through the initial mixtures. As scientists better understand the characteristics of combustion synthesis, they can refine the technique to be more useful in advanced materials production.

Questions

1. Why is this technique called *combustion synthesis?*
2. Why might this technique result in a more chemically homogeneous material?

▼ *Once the reactant mixture is ignited, the combustion wave moves through the sample, synthesizing the solid-state product.*



CHAPTER HIGHLIGHTS

Describing Chemical Reactions

Vocabulary
chemical equation
precipitate
coefficient
word equation
formula equation
reversible reaction

- Four observations that suggest a chemical reaction is taking place are the evolution of energy as heat and 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.

Types of Chemical Reactions

Vocabulary
synthesis reaction
decomposition reaction
electrolysis
single-displacement reaction
double-displacement reaction
combustion reaction

- 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-displacement reactions are represented by the general equations $A + BX \longrightarrow AX + B$ and $Y + BX \longrightarrow BY + X$.
- Double-displacement reactions are represented by the general equation $AX + BY \longrightarrow AY + BX$.
- In a combustion reaction, a substance combines with oxygen, releasing energy in the form of heat and light.

Activity Series of the Elements

Vocabulary
activity series

- 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.

CHAPTER REVIEW

Describing Chemical Reactions

SECTION 1 REVIEW

1. List four observations that indicate that a chemical reaction may be taking place.

2. List the three requirements for a correctly written chemical equation.

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 the coefficient precedes?

4. Give an example of a word equation, a formula equation, and a chemical equation.

5. What quantitative information is revealed by a chemical equation?

6. What limitations are associated with the use of both word and formula equations?

7. Define each of the following terms:
a. aqueous solution
b. catalyst
c. reversible reaction

8. Write formulas for each of the following compounds:
a. potassium hydroxide
b. calcium nitrate
c. sodium carbonate
d. carbon tetrachloride
e. magnesium bromide

9. What four guidelines are useful in balancing an equation?

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. $4C_3H_8$

PRACTICE PROBLEMS

11. Write the chemical equation that relates to each of the following word equations. Include symbols for physical states in the equation.
a. solid zinc sulfide + oxygen gas $\longrightarrow$ solid zinc oxide + sulfur dioxide gas
b. aqueous hydrochloric acid + aqueous barium hydroxide $\longrightarrow$ aqueous barium chloride + water
c. aqueous nitric acid + aqueous calcium hydroxide $\longrightarrow$ aqueous calcium nitrate + water

12. Translate each of the following chemical equations into a sentence.
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. $AgNO_3(aq) + KI(aq) \longrightarrow AgI(s) + KNO_3(aq)$

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

14. Identify and correct each error in the following equations, 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. $NaI + Cl_2 \longrightarrow NaCl + I$

15. 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.

16. 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 C and D.)

17. 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

SECTION 2 REVIEW

18. Define and give general equations for the five basic types of chemical reactions introduced in Chapter 8.
19. How are most decomposition reactions initiated?
20. A substance is decomposed by an electric current. What is the name of this type of reaction?
21. a. In what environment do many single-displacement reactions commonly occur?
    b. In general, how do single-displacement reactions compare with synthesis and decomposition reactions in terms of the amount of energy involved?

PRACTICE PROBLEMS

22. Complete each of the following synthesis reactions by writing both a word equation and a chemical equation.
    a. sodium + oxygen $\longrightarrow$ ____
    b. magnesium + fluorine $\longrightarrow$ ____
23. Complete and balance the equations for the following decomposition reactions:
    a. $HgO \xrightarrow{\Delta}$
    b. $H_2O(l) \xrightarrow{\text{electricity}}$
    c. $Ag_2O \xrightarrow{\Delta}$
    d. $CuCl_2 \xrightarrow{\text{electricity}}$
24. Complete and balance the equations for the following single-displacement reactions:
    a. $Zn + Pb(NO_3)_2 \longrightarrow \_\_\_\_$
    b. $Al + Hg(CH_3COO)_2 \longrightarrow \_\_\_\_$
    c. $Al + NiSO_4 \longrightarrow \_\_\_\_$
    d. $Na + H_2O \longrightarrow \_\_\_\_$
25. Complete and balance the equations for the following double-displacement reactions:
    a. $AgNO_3(aq) + NaCl(aq) \longrightarrow \_\_\_\_$
    b. $Mg(NO_3)_2(aq) + KOH(aq) \longrightarrow \_\_\_\_$
    c. $LiOH(aq) + Fe(NO_3)_3(aq) \longrightarrow \_\_\_\_$
26. Complete and balance the equations for the following combustion reactions:
    a. $CH_4 + O_2 \longrightarrow \_\_\_\_$
    b. $C_3H_6 + O_2 \longrightarrow \_\_\_\_$
    c. $C_5H_{12} + O_2 \longrightarrow \_\_\_\_$
27. 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
28. 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
29. 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. $C_2H_5OH + \_\_\_\_ \longrightarrow \_\_\_\_ + \_\_\_\_$
30. Complete and balance the equations for the following reactions, and then identify each by type.
    a. zinc + sulfur $\longrightarrow$ ____
    b. silver nitrate + potassium iodide $\longrightarrow$ ____
    c. toluene, C_7H_8 + oxygen $\longrightarrow$ ____
    d. nonane, C_9H_{20} + oxygen $\longrightarrow$ ____

Activity Series of the Elements

SECTION 3 REVIEW

31. a. What is meant by the *activity* of an element?
    b. How does this description differ for metals and nonmetals?

32. a. What is an activity series of elements?
b. What is the basis for the ordering of the elements in the activity series?

33. a. What chemical principle is the basis for the activity series of metals?
b. What is the significance of the distance between two metals in the activity series?

PRACTICE PROBLEMS

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 3** on page 286, 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. $Pb(s) + ZnCl_2(s) \longrightarrow$ ______

43. Complete and balance the equations for the following reactions, and identify the type of reaction that each equation 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 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. $Al(OH)_3(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2O(l)$

CRITICAL THINKING

48. **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. Using your knowledge of electron configurations and periodic trends, infer possible explanations for the metals' reactivity and position in the periodic table.

49. **Analyzing Results** Formulate an activity series for the hypothetical elements A, J, Q, and Z by using the following reaction information:

$$A + ZX \longrightarrow AX + Z$$

$$J + ZX \longrightarrow \text{no reaction}$$

$$Q + AX \longrightarrow QX + A$$

USING THE HANDBOOK

50. 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).

51. 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).

RESEARCH & WRITING

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

53. 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

54. **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.

Math Tutor BALANCING CHEMICAL EQUATIONS

A chemical equation is a written expression of an actual chemical reaction in which certain atoms, ions, or molecules become rearranged in a specific way. Therefore, the equation must represent exactly what happens in the reaction. Recall that atoms are never created or destroyed in chemical reactions. A balanced chemical equation shows that all of the atoms present in reactants are still present in products.

Problem-Solving TIPS

- First, identify reactants and products. (You may find it helpful to write a word equation first.)
- Using correct formulas and symbols, write an unbalanced equation for the reaction.
- Balance atoms one element at a time by inserting coefficients.
- Identify elements that appear in only one reactant and one product, and balance the atoms of those elements first.
- If a polyatomic ion appears on both sides of the equation, treat it as a single unit.
- Double-check to be sure that the number of atoms of each element is the same on both sides of the equation.

SAMPLE

When an aqueous solution of ammonium sulfate, $(NH_4)_2SO_4(aq)$, is combined with an aqueous solution of silver nitrate, $AgNO_3(aq)$, a precipitate of solid silver sulfate, $Ag_2SO_4(s)$, forms, leaving ammonium nitrate, $NH_4NO_3(aq)$, in solution. Balance the equation for this reaction.

As before, first write an equation with correct formulas for all reactants and products.

$$(NH_4)_2SO_4(aq) + AgNO_3(aq) \longrightarrow NH_4NO_3(aq) + Ag_2SO_4(s)$$

If you compare the number of silver atoms on each side, you can see that the equation is not balanced. This equation may look very complex, but it is really fairly simple. In many reactions involving polyatomic ions such as sulfate, nitrate, and ammonium, the ions do not change. In the equation above, you can see that NO_3 is present on both sides, as are SO_4 and NH_4. You can balance the equation by treating the groups as if they were single atoms. To balance the NH_4 groups, place a 2 in front of NH_4NO_3. This gives you two ammonium groups on the left and two on the right. Now, because you have two nitrate groups on the right, place a 2 in front of $AgNO_3$ to give two nitrate groups on the left. Finally, check silver atoms and sulfate groups, and you find that they balance.

$$(NH_4)_2SO_4(aq) + 2AgNO_3(aq) \longrightarrow 2NH_4NO_3(aq) + Ag_2SO_4(s)$$

PRACTICE PROBLEMS

1. When propane burns completely in air, the reaction forms carbon dioxide and water vapor. Balance the equation for this reaction.

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

2. Balance the following chemical equations:
   a. $KI(aq) + Cl_2(g) \longrightarrow KCl(aq) + I_2(s)$
   b. $Al(s) + H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + H_2(g)$

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. According to the law of conservation of mass, the total mass of the reacting substances is
A. always more than the total mass of the products.
B. always less than the total mass of the products.
C. sometimes more and sometimes less than the total mass of the products.
D. always equal to the total mass of the products.

2. To balance a chemical equation, you may adjust the
A. coefficients.
B. subscripts.
C. formulas of the products.
D. either the coefficients or the subscripts.

3. Which is the correct chemical equation for the following formula equation: $(NH_4)_2S \longrightarrow NH_3 + H_2S$?
A. $2(NH_4)_2S \longrightarrow 2NH_3 + H_2S_2$
B. $2(NH_4)_2S \longrightarrow 2NH_3 + H_2S$
C. $(NH_4)_2S \longrightarrow 2NH_3 + H_2S$
D. None of the above

4. Select the missing reactant(s) for the double-displacement reaction that produces PF_5 and $AsCl_3$.
A. PCl_5 and AsF_3
B. PCl_3 and AsF_5
C. PCl_3 and AsF_3
D. None of the above

5. Select the missing reactant for the following combustion reaction: $2\_\_\_\_\_\_ + 15O_2 \longrightarrow 14CO_2 + 6H_2O$.
A. $C_{14}H_{12}$
B. $C_{14}H_{12}O_4$
C. C_7H_6
D. $C_7H_6O_2$

6. A mixture consists of Ag, Pb, and Fe metals. Which of these metals will react with $ZnCl_2$?
A. Ag(s)
B. Pb(s)
C. Fe(s)
D. None of these metals

7. Which of the following statements is true about the reaction $2F_2 + 2H_2O \longrightarrow 4HF + O_2$?
A. Two grams of O_2 are produced when 2 g F_2 reacts with 2 g H_2O.
B. Two moles of HF are produced when 1 mol F_2 reacts with 1 mol H_2O.
C. For every 2 mol O_2 produced, 6 mol HF are produced.
D. For every 1 mol H_2O that reacts, 2 mol O_2 are produced.

SHORT ANSWER

8. Determine the products and write a balanced equation for the reaction of solid magnesium and water.

9. A precipitation of iron(III) hydroxide is produced by reacting an aqueous solution of iron(III) chloride with an aqueous solution of sodium hydroxide. Write a balanced chemical equation.

EXTENDED RESPONSE

10. List the hypothetical metals A, E, M, and R in increasing order of reactivity by using the reaction data in the table below. The reaction of interest is of the form $C + ZX \longrightarrow CX + Z$. Explain your reasoning.

	AX	EX	MX	RX
A	——	no reaction	reaction	no reaction
E	reaction	——	reaction	reaction
M	no reaction	no reaction	——	no reaction
R	reaction	no reaction	reaction	——

11. Calcium hypochlorite, $Ca(OCl)_2$, is a bleaching agent produced from sodium hydroxide, calcium hydroxide, and chlorine. Sodium chloride and water are also produced in the reaction. Write the balanced chemical equation. If 2 mol NaOH react, how many moles of calcium hypochlorite can be produced?

Test TIP Focus on one question at a time unless you are asked to refer to previous answers.

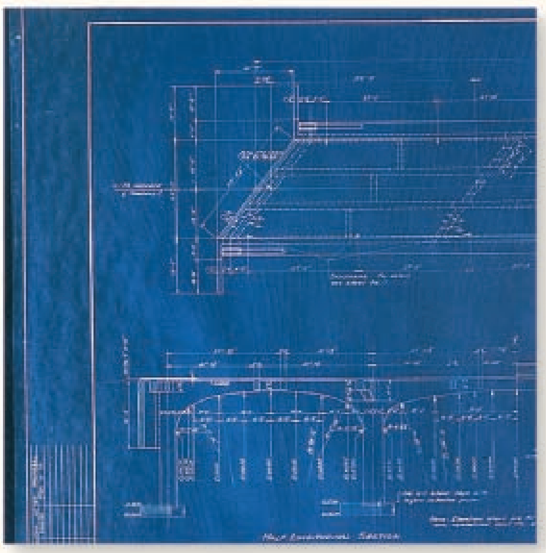
Blueprint Paper

OBJECTIVES

- ***Prepare* blueprint paper and create a blueprint.**

MATERIALS

- **10% iron(III) ammonium citrate solution**
- **10% potassium hexacyanoferrate(III) solution**
- **25 mL graduated cylinders, 2**
- **corrugated cardboard, 20 cm × 30 cm, 2 pieces**
- **glass stirring rod**
- **Petri dish**
- **thumbtacks, 4**
- **tongs**
- **white paper, 8 cm × 15 cm, 1 piece**



BACKGROUND

Blueprint paper is prepared by coating paper with a solution of two soluble iron(III) salts—potassium hexacyanoferrate(III), commonly called *potassium ferricyanide,* and iron(III) ammonium citrate. These two salts do not react with each other in the dark. However, when exposed to UV light, the iron(III) ammonium citrate is converted to an iron(II) salt. Potassium hexacyanoferrate(III), $K_3Fe(CN)_6$, reacts with iron(II) ion, Fe^{2+}, to produce an insoluble blue compound, $KFeFe(CN)_6{\bullet}H_2O$. In this compound, iron appears to exist in both the +2 and +3 oxidation states.

A blueprint is made by using black ink to make a sketch on a piece of tracing paper or clear, colorless plastic. This sketch is placed on top of a piece of blueprint paper and exposed to ultraviolet light. Wherever the light strikes the paper, the paper turns blue. The paper is then washed to remove the soluble unexposed chemical and is allowed to dry. The result is a blueprint—a blue sheet of paper with white lines.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PROCEDURE

1. Pour 15 mL of a 10% solution of potassium hexacyanoferrate(III) solution into a Petri dish. With most of the classroom lights off or dimmed, add 15 mL of 10% iron(III) ammonium citrate solution. Stir the mixture.

2. Write your name on an 8 cm × 15 cm piece of white paper. Carefully coat one side of the piece of paper by using tongs to drag it over the top of the solution in the Petri dish.

3. With the coated side up, tack your wet paper to a piece of corrugated cardboard, and cover the paper with another piece of cardboard. **Wash your hands before proceeding to step 4.**

4. Take your paper and cardboard assembly outside into the direct sunlight. Remove the top piece of cardboard so that the paper is exposed. Quickly place an object such as a fern, a leaf, or a key on the paper. If it is windy, you may need to put small weights, such as coins, on the object to keep it in place, as shown in **Figure A.**

5. After about 20 min, remove the object and again cover the paper with the cardboard. Return to the lab, remove the tacks, and *thoroughly* rinse the blueprint paper under cold running water. Allow the paper to dry. In your notebook, record the amount of time that the paper was exposed to sunlight.

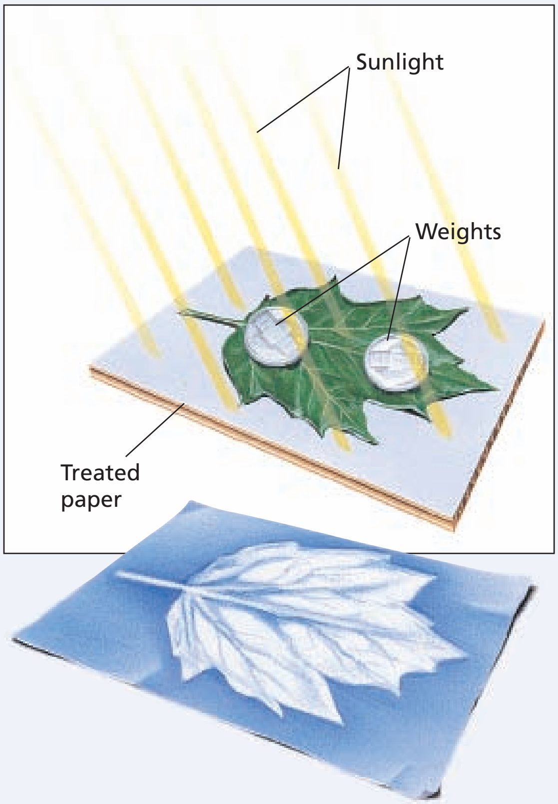


FIGURE A To produce a sharp image, the object must be flat, with its edges on the blueprint paper, and it must not move.

CLEANUP AND DISPOSAL

6. Clean all equipment and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. **Relating Ideas:** Why is the iron(III) ammonium citrate solution stored in a brown bottle?

2. **Organizing Ideas:** When iron(III) ammonium citrate is exposed to light, the oxidation state of the iron changes. What is the new oxidation state of the iron?

3. **Analyzing Methods:** What substances were washed away when you rinsed the blueprint in water after it had been exposed to sunlight? (Hint: Compare the solubilities of the two ammonium salts that you used to coat the paper and of the blue product that formed.)

CONCLUSIONS

1. **Applying Ideas:** Insufficient washing of the exposed blueprints results in a slow deterioration of images. Suggest a reason for this deterioration.

2. **Relating Ideas:** Photographic paper can be safely exposed to red light in a darkroom. Do you think the same would be true of blueprint paper? Explain your answer.

EXTENSIONS

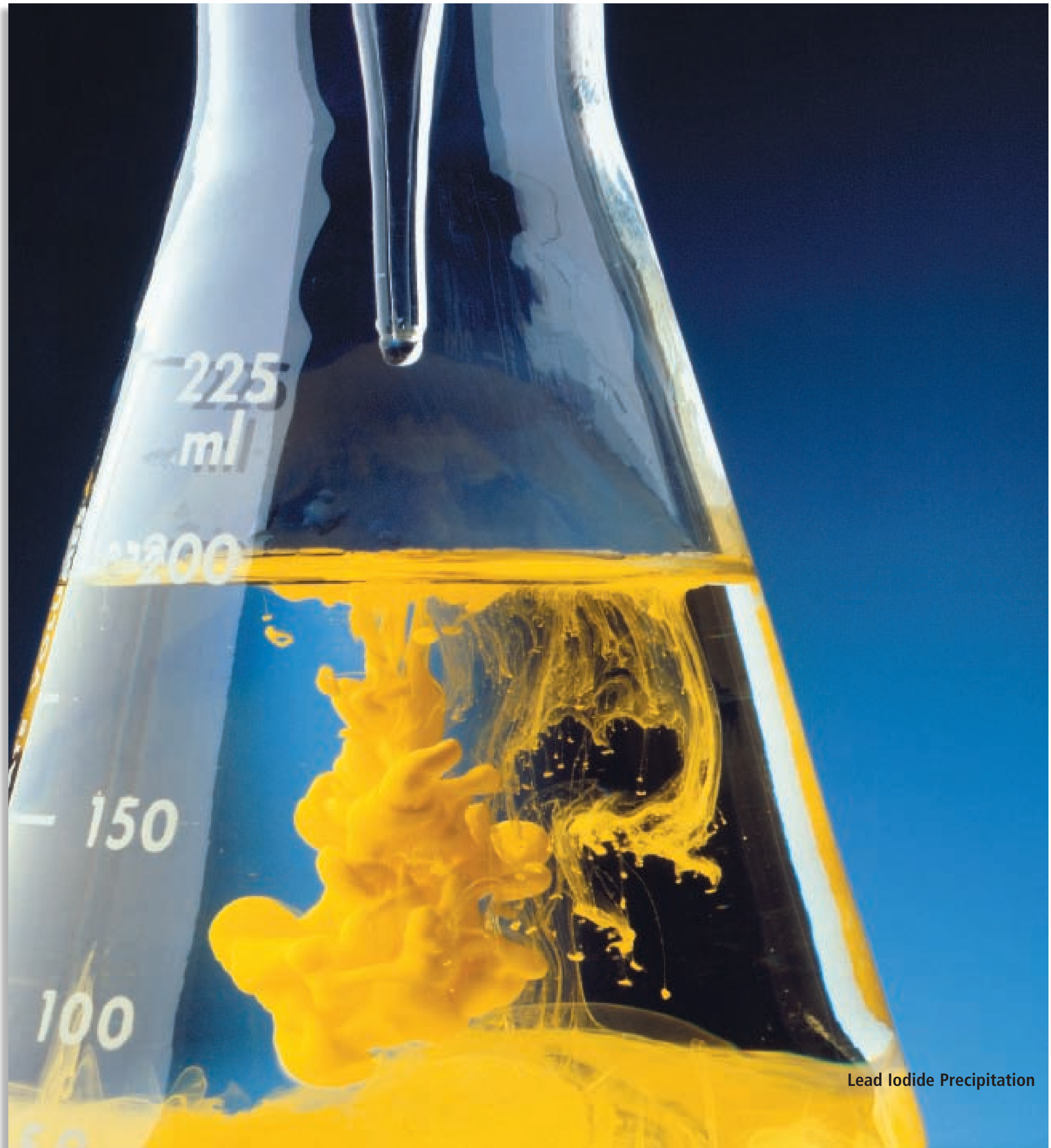
1. **Applying Ideas:** How could you use this blueprint paper to test the effectiveness of a brand of sunscreen lotion?

2. **Designing Experiments:** Can you think of ways to improve this procedure? If so, ask your teacher to approve your plan, and create a new blueprint. Evaluate both the efficiency of the procedure and the quality of blueprint.

CHAPTER 9

Stoichiometry

Stoichiometry *comes from the Greek words* stoicheion, *meaning "element," and* metron, *meaning "measure."*



Lead Iodide Precipitation

Introduction to Stoichiometry

SECTION 1

OBJECTIVES

- Define *stoichiometry.*
- Describe the importance of the *mole ratio* in stoichiometric calculations.
- Write a mole ratio relating two substances in a chemical equation.

Much of our knowledge of chemistry is based on the careful quantitative analysis of substances involved in chemical reactions. **Composition stoichiometry** (which you studied in Chapter 3) *deals with the mass relationships of elements in compounds*. **Reaction stoichiometry** *involves the mass relationships between reactants and products in a chemical reaction.* Reaction stoichiometry is the subject of this chapter and it is based on chemical equations and the law of conservation of mass. All reaction stoichiometry calculations start with a balanced chemical equation. This equation gives the relative numbers of moles of reactants and products.

Reaction Stoichiometry Problems

The reaction stoichiometry problems in this chapter can be classified according to the information *given* in the problem and the information you are expected to find, the *unknown*. The *given* and the *unknown* may both be reactants, they may both be products, or one may be a reactant and the other a product. The masses are generally expressed in grams, but you will encounter both large-scale and microscale problems with other mass units, such as kg or mg. Stoichiometric problems are solved by using ratios from the balanced equation to convert the given quantity using the methods described here.

Problem Type 1: *Given* and *unknown* quantities are amounts in moles.
When you are given the amount of a substance in moles and asked to calculate the amount in moles of another substance in the chemical reaction, the general plan is

$$\begin{array}{c}\text{amount of}\\ \textit{given}\text{ substance (mol)}\end{array} \longrightarrow \begin{array}{c}\text{amount of}\\ \textit{unknown}\text{ substance (mol)}\end{array}$$

Problem Type 2: *Given* is an amount in moles and *unknown* is a mass that is often expressed in grams.
When you are given the amount in moles of one substance and asked to calculate the mass of another substance in the chemical reaction, the general plan is

$$\begin{array}{c}\text{amount of}\\ \textit{given}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{amount of}\\ \textit{unknown}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{mass of}\\ \textit{unknown}\text{ substance}\\ \text{(g)}\end{array}$$

Problem Type 3: *Given* is a mass in grams and *unknown* is an amount in moles.

When you are given the mass of one substance and asked to calculate the amount in moles of another substance in the chemical reaction, the general plan is

$$\underset{(\text{g})}{\text{mass of } \textit{given} \text{ substance}} \longrightarrow \underset{(\text{mol})}{\text{amount of } \textit{given} \text{ substance}} \longrightarrow \underset{(\text{mol})}{\text{amount of } \textit{unknown} \text{ substance}}$$

Problem Type 4: *Given* is a mass in grams and *unknown* is a mass in grams.

When you are given the mass of one substance and asked to calculate the mass of another substance in the chemical reaction, the general plan is

$$\underset{(\text{g})}{\text{mass of } \textit{given} \text{ substance}} \longrightarrow \underset{(\text{mol})}{\text{amount of } \textit{given} \text{ substance}} \longrightarrow \underset{(\text{mol})}{\text{amount of } \textit{unknown} \text{ substance}} \longrightarrow \underset{(\text{g})}{\text{mass of } \textit{unknown} \text{ substance}}$$

Mole Ratio

Solving any reaction stoichiometry problem requires the use of a mole ratio to convert from moles or grams of one substance in a reaction to moles or grams of another substance. *A* **mole ratio** *is a conversion factor that relates the amounts in moles of any two substances involved in a chemical reaction.* This information is obtained directly from the balanced chemical equation. Consider, for example, the chemical equation for the electrolysis of melted aluminum oxide to produce aluminum and oxygen.

$$2Al_2O_3(l) \longrightarrow 4Al(s) + 3O_2(g)$$

Recall from Chapter 8 that the coefficients in a chemical equation satisfy the law of conservation of matter and represent the relative amounts in moles of reactants and products. Therefore, 2 mol of aluminum oxide decompose to produce 4 mol of aluminum and 3 mol of oxygen gas. These relationships can be expressed in the following mole ratios.

$$\frac{2\text{ mol }Al_2O_3}{4\text{ mol Al}} \text{ or } \frac{4\text{ mol Al}}{2\text{ mol }Al_2O_3}$$

$$\frac{2\text{ mol }Al_2O_3}{3\text{ mol }O_2} \text{ or } \frac{3\text{ mol }O_2}{2\text{ mol }Al_2O_3}$$

$$\frac{4\text{ mol Al}}{3\text{ mol }O_2} \text{ or } \frac{3\text{ mol }O_2}{4\text{ mol Al}}$$

For the decomposition of aluminum oxide, the appropriate mole ratio would be used as a conversion factor to convert a given amount in moles of one substance to the corresponding amount in moles of another

CAREERS in Chemistry

Chemical Technician

Chemical technicians are highly skilled scientific professionals who bring valuable skills to the development of new products, the processing of materials, the management of hazardous waste, regulatory compliance, and many other aspects of getting products and services to the consumer. Chemical technicians must have a solid background in applied chemistry and mathematics and be highly skilled in laboratory methods. Earning an associate's degree in applied science or chemical technology is one good way to prepare for this career. Many chemical technicians have a bachelor's degree in chemical technology, chemistry, or other sciences.

substance. To determine the amount in moles of aluminum that can be produced from 13.0 mol of aluminum oxide, the mole ratio needed is that of Al to Al_2O_3.

$$13.0\ \cancel{\text{mol } Al_2O_3} \times \frac{4 \text{ mol Al}}{2\ \cancel{\text{mol } Al_2O_3}} = 26.0 \text{ mol Al}$$

Mole ratios are exact, so they do not limit the number of significant figures in a calculation. The number of significant figures in the answer is therefore determined only by the number of significant figures of any measured quantities in a particular problem.

extension

Chemistry in Action

Go to **go.hrw.com** for a full-length article on stoichiometry and air bags.

Keyword: HC6STCX

Molar Mass

Recall from Chapter 7 that the molar mass is the mass, in grams, of one mole of a substance. The molar mass is the conversion factor that relates the mass of a substance to the amount in moles of that substance. To solve reaction stoichiometry problems, you will need to determine molar masses using the periodic table.

Returning to the previous example, the decomposition of aluminum oxide, the rounded masses from the periodic table are the following.

$$1 \text{ mol } Al_2O_3 = 101.96 \text{ g} \qquad 1 \text{ mol Al} = 26.98 \text{ g} \qquad 1 \text{ mol } O_2 = 32.00 \text{ g}$$

These molar masses can be expressed by the following conversion factors.

$$\frac{101.96 \text{ g } Al_2O_3}{1 \text{ mol } Al_2O_3} \text{ or } \frac{1 \text{ mol } Al_2O_3}{101.96 \text{ g } Al_2O_3}$$

$$\frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \text{ or } \frac{1 \text{ mol Al}}{26.98 \text{ g Al}}$$

$$\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} \text{ or } \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2}$$

To find the number of grams of aluminum equivalent to 26.0 mol of aluminum, the calculation would be as follows.

$$26.0\ \cancel{\text{mol Al}} \times \frac{26.98 \text{ g Al}}{1\ \cancel{\text{mol Al}}} = 701 \text{ g Al}$$

SECTION REVIEW

1. What is stoichiometry?
2. For each equation, write all possible mole ratios.
   a. $2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$
   b. $4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(l)$
3. How is a mole ratio used in stoichiometry?

Critical Thinking

4. **RELATING IDEAS** What step must be performed before any stoichiometry problem is solved? Explain.

HISTORICAL CHEMISTRY

The Case of Combustion

People throughout history have transformed substances by burning them in air. Yet at the dawn of the scientific revolution, very little was known about the process of combustion. In attempting to explain this common phenomenon, chemists of the 18th century developed one of the first universally accepted theories in their field. But as one man would show, scientific theories do not always stand the test of time.

▲*Antoine Laurent Lavoisier helped establish chemistry as a science.*

Changing Attitudes

Shunning the ancient Greek approach of logical argument based on untested premises, investigators of the 17th century began to understand the laws of nature by observing, measuring, and performing experiments on the world around them. However, this scientific method was incorporated into chemistry slowly. Although early chemists experimented extensively, most considered measurement to be unimportant. This viewpoint hindered the progress of chemistry for nearly a century.

A Flawed Theory

By 1700, combustion was assumed to be the decomposition of a material into simpler substances. People saw burning substances emitting smoke and energy as heat and light. To account for this emission, scientists proposed a theory that combustion depended on the emission of a substance called *phlogiston,* which appeared as a combination of energy as heat and light while the material was burning but which could not be detected beforehand.

The phlogiston theory was used to explain many chemical observations of the day. For example, a lit candle under a glass jar burned until the surrounding air became saturated with phlogiston, at which time the flame died because the air inside could not absorb more phlogiston.

A New Phase of Study

By the 1770s, the phlogiston theory had gained universal acceptance. At that time, chemists also began to experiment with air, which was generally believed to be an element.

In 1772, when Daniel Rutherford found that a mouse kept in a closed container soon died, he explained the results based on the phlogiston theory. Like a burning candle, the mouse emitted phlogiston; when the air could hold no more phlogiston, the mouse died. Thus, Rutherford figured that the air in the container had become "phlogisticated air."

A couple of years later, Joseph Priestley obtained a reddish powder when he heated mercury in the air. He assumed that the powder was mercury devoid of phlogiston. But when he heated the powder, an unexpected result occurred: Metallic mercury, along with a gas that allowed a candle to burn, formed. Following the phlogiston theory, he believed this gas that supports combustion to be "dephlogisticated air."

Nice Try, But . . .

Antoine Laurent Lavoisier was a meticulous scientist. He realized that Rutherford and Priestley had carefully observed and described their experiments but had not measured the mass of anything. Unlike his colleagues, Lavoisier knew the importance of using a balance. He measured the masses of reactants and products and compared them. He observed that the total mass of the reactants equaled the total mass of the products. Based on these observations, which supported what would become known as the *law of conservation of mass,* Lavoisier endeavored to explain the results of Rutherford and Priestley.

Lavoisier put some tin in a closed vessel and weighed the entire system. He then burned the tin. When he opened the vessel, air rushed into it as if something had been *removed* from the air in the vessel during combustion. He then measured the mass of the burned metal and observed that this mass was greater than the mass of the original tin. Curiously, this increase in mass equaled the mass of the air that had rushed into the vessel. To Lavoisier, this change in mass did not support the idea of phlogiston escaping the burning material. Instead, it indicated that during combustion, part of the air reacted with the tin.

After obtaining similar results by using various substances, Lavoisier concluded that air was not an element but a mixture composed principally of two gases, Priestley's "dephlogisticated air" (which Lavoisier renamed *oxygen*) and Rutherford's "phlogisticated air" (which was mostly nitrogen but had traces of other nonflammable atmospheric gases). When a substance burned, it chemically combined with oxygen, resulting in a product Lavoisier named an *oxide.* Lavoisier's theory of combustion persists today. He used the name *oxygen* because he thought that all acids contained oxygen. *Oxygen* means "acid former."

The Father of Chemistry

By emphasizing the importance of quantitative analysis, Lavoisier helped establish chemistry as a science. His work on combustion laid to rest the phlogiston theory and the theory that air is an element. He also explained why hydrogen burned in oxygen to form water, or hydrogen oxide. He later published one of the first chemistry textbooks, which established a common naming system of compounds and elements and helped unify chemistry worldwide. These accomplishments earned Lavoisier the reputation of being the father of chemistry.

Questions

1. Why does the mass of tin increase when tin is heated in air?
2. What was the composition of Priestley's "dephlogisticated air" and Rutherford's "phlogisticated air"?

TABLE OF SIMPLE SUBSTANCES.

Simple fubftances belonging to all the kingdoms of nature, which may be confidered as the elements of bodies.

New Names.	*Correfpondent old Names.*
Light - - -	*Light.*
Caloric - - -	*Heat.* *Principle or element of heat.* *Fire. Igneous fluid.* *Matter of fire and of heat.*
Oxygen - - -	*Dephlogifticated air.* *Empyreal air.* *Vital air, or* *Bafe of vital air.*
Azote - - -	*Phlogifticated air or gas.* *Mephitis, or its bafe.*
Hydrogen - -	*Inflammable air or gas,* *or the bafe of inflammable air.*

▲ *Lavoisier's concept of simple substances was published in his book* Elements of Chemistry *in 1789.*

SECTION 2

Ideal Stoichiometric Calculations

Objectives

- Calculate the amount in moles of a reactant or product from the amount in moles of a different reactant or product.
- Calculate the mass of a reactant or product from the amount in moles of a different reactant or product.
- Calculate the amount in moles of a reactant or product from the mass of a different reactant or product.
- Calculate the mass of a reactant or product from the mass of a different reactant or product.

The balanced chemical equation is the key step in all stoichiometric calculations because the mole ratio is obtained directly from it. Solving any reaction stoichiometry problem must begin with a balanced equation.

Chemical equations help us plan the amounts of reactants to use in a chemical reaction without having to run the reactions in the laboratory. The reaction stoichiometry calculations described in this chapter are theoretical. They tell us the amounts of reactants and products for a given chemical reaction under *ideal conditions,* in which all reactants are completely converted into products. However, many reactions do not proceed such that all reactants are completely converted into products. Theoretical stoichiometric calculations allow us to determine the maximum amount of product that could be obtained in a reaction when the reactants are not pure or byproducts are formed in addition to the expected products.

Solving stoichiometric problems requires practice. These problems are extensions of the composition stoichiometry problems that you solved in Chapters 3 and 7. Practice by working the sample problems in the rest of this chapter. Using a logical, systematic approach will help you successfully solve these problems.

Conversions of Quantities in Moles

In these stoichiometric problems, you are asked to calculate the amount in moles of one substance that will react with or be produced from the given amount in moles of another substance. The plan for a simple mole conversion problem is

$$\begin{matrix}\text{amount of} \\ \textit{given}\text{ substance (mol)}\end{matrix} \longrightarrow \begin{matrix}\text{amount of} \\ \textit{unknown}\text{ substance (mol)}\end{matrix}$$

This plan requires one conversion factor—the stoichiometric mole ratio of the *unknown* substance to the *given* substance from the balanced equation. To solve this type of problem, simply multiply the *known* quantity by the appropriate conversion factor.

$$\textit{given}\text{ quantity} \times \text{conversion factor} = \textit{unknown}\text{ quantity}$$

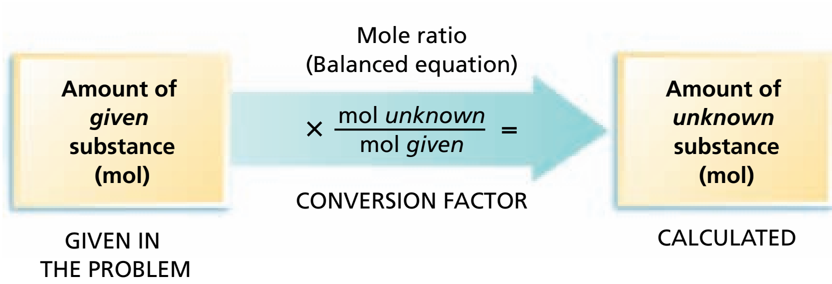


FIGURE 1 This is a solution plan for problems in which the given and unknown quantities are expressed in moles.

SAMPLE PROBLEM A

For more help, go to the ***Math Tutor*** *at the end of this chapter.*

In a spacecraft, the carbon dioxide exhaled by astronauts can be removed by its reaction with lithium hydroxide, LiOH, according to the following chemical equation.

$$CO_2(g) + 2LiOH(s) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

How many moles of lithium hydroxide are required to react with 20 mol CO_2, the average amount exhaled by a person each day?

SOLUTION

1 *ANALYZE*

Given: amount of CO_2 = 20 mol
Unknown: amount of LiOH (mol)

2 *PLAN*

amount of CO_2 (mol) $\longrightarrow$ amount of LiOH (mol)

This problem requires one conversion factor—the mole ratio of LiOH to CO_2. The mole ratio is obtained from the balanced chemical equation. Because you are given moles of CO_2, select a mole ratio that will cancel mol CO_2 and give you mol LiOH in your final answer. The correct ratio has the following units.

$$\frac{\text{mol LiOH}}{\text{mol } CO_2}$$

This ratio cancels mol CO_2 and gives the units mol LiOH in the answer.

$$\text{mol } CO_2 \times \overbrace{\frac{\text{mol LiOH}}{\text{mol } CO_2}}^{\text{mol ratio}} = \text{mol LiOH}$$

3 *COMPUTE* Substitute the values in the equation in step 2, and compute the answer.

$$20\ \cancel{\text{mol } CO_2} \times \frac{2 \text{ mol LiOH}}{1\ \cancel{\text{mol } CO_2}} = 40 \text{ mol LiOH}$$

4 *EVALUATE* The answer is written correctly with one significant figure to match the number of significant figures in the given 20 mol CO_2, and the units cancel to leave mol LiOH, which is the unknown. The balanced equation shows that twice the amount in moles of LiOH reacts with CO_2. Therefore, the answer should be $2 \times 20 = 40$.

PRACTICE

Answers in Appendix E

1. Ammonia, NH_3, is widely used as a fertilizer and in many household cleaners. How many moles of ammonia are produced when 6 mol of hydrogen gas react with an excess of nitrogen gas?
2. The decomposition of potassium chlorate, $KClO_3$, into KCl and O_2 is used as a source of oxygen in the laboratory. How many moles of potassium chlorate are needed to produce 15 mol of oxygen gas?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate unknown quantities by using mole ratios.

Keyword: HC6STCX

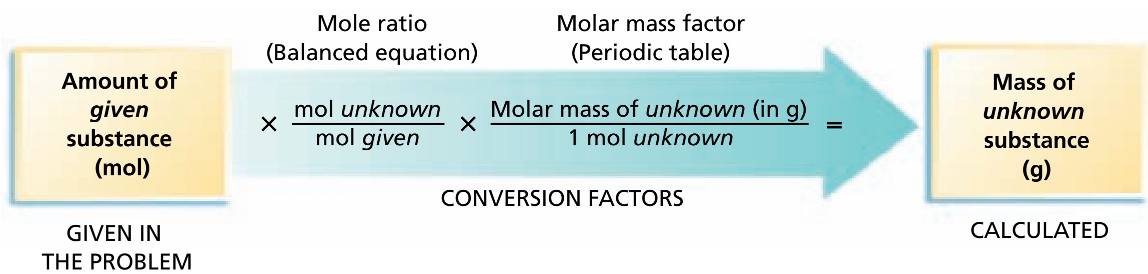
Conversions of Amounts in Moles to Mass

In these stoichiometric calculations, you are asked to calculate the mass (usually in grams) of a substance that will react with or be produced from a given amount in moles of a second substance. The plan for these mole-to-gram conversions is

$$\begin{array}{c}\text{amount of}\\ \textit{given}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{amount of}\\ \textit{unknown}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{mass of}\\ \textit{unknown}\text{ substance}\\ \text{(g)}\end{array}$$

This plan requires two conversion factors—the mole ratio of the *unknown* substance to the *given* substance and the molar mass of the *unknown* substance for the mass conversion. To solve this kind of problem, you simply multiply the known quantity, which is the amount in moles, by the appropriate conversion factors.

FIGURE 2 This is a solution plan for problems in which the given quantity is expressed in moles and the unknown quantity is expressed in grams.



SAMPLE PROBLEM B

For more help, go to the Math Tutor at the end of this chapter.

In photosynthesis, plants use energy from the sun to produce glucose, $C_6H_{12}O_6$, and oxygen from the reaction of carbon dioxide and water. What mass, in grams, of glucose is produced when 3.00 mol of water react with carbon dioxide?

SOLUTION

1 ***ANALYZE***

Given: amount of H_2O = 3.00 mol

Unknown: mass of $C_6H_{12}O_6$ produced (g)

2 *PLAN* You must start with a balanced equation.

$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(s) + 6O_2(g)$$

Given the amount in mol of H_2O, you need to get the mass of $C_6H_{12}O_6$ in grams. Two conversion factors are needed—the mole ratio of $C_6H_{12}O_6$ to H_2O and the molar mass of $C_6H_{12}O_6$.

$$\text{mol } H_2O \times \overset{\text{mol ratio}}{\frac{\text{mol } C_6H_{12}O_6}{\text{mol } H_2O}} \times \overset{\text{molar mass factor}}{\frac{\text{g } C_6H_{12}O_6}{\text{mol } C_6H_{12}O_6}} = \text{g } C_6H_{12}O_6$$

3 *COMPUTE* Use the periodic table to compute the molar mass of $C_6H_{12}O_6$.

$$C_6H_{12}O_6 = 180.18 \text{ g/mol}$$

$$3.00 \cancel{\text{mol } H_2O} \times \frac{1 \cancel{\text{mol } C_6H_{12}O_6}}{6 \cancel{\text{mol } H_2O}} \times \frac{180.18 \text{ g } C_6H_{12}O_6}{1 \cancel{\text{mol } C_6H_{12}O_6}} = 90.1 \text{ g } C_6H_{12}O_6$$

4 *EVALUATE* The answer is correctly rounded to three significant figures, to match those in 3.00 mol H_2O. The units cancel in the problem, leaving g $C_6H_{12}O_6$ as the units for the answer, which matches the unknown. The answer is reasonable because it is one-half of 180.

SAMPLE PROBLEM C

For more help, go to the Math Tutor at the end of this chapter.

What mass of carbon dioxide, in grams, is needed to react with 3.00 mol H_2O in the photosynthetic reaction described in Sample Problem B?

SOLUTION

1 *ANALYZE*

Given: amount of H_2O = 3.00 mol

Unknown: mass of CO_2 (g)

2 *PLAN* The chemical equation from Sample Problem B is

$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(s) + 6O_2(g).$$

Two conversion factors are needed—the mole ratio of CO_2 to H_2O and the molar mass factor of CO_2.

$$\text{mol } H_2O \times \overset{\text{mol ratio}}{\frac{\text{mol } CO_2}{\text{mol } H_2O}} \times \overset{\text{molar mass factor}}{\frac{\text{g } CO_2}{\text{mol } CO_2}} = \text{g } CO_2$$

3 *COMPUTE* Use the periodic table to compute the molar mass of CO_2.

$$CO_2 = 44.01 \text{ g/mol}$$

$$3.00 \cancel{\text{mol } H_2O} \times \frac{6 \cancel{\text{mol } CO_2}}{6 \cancel{\text{mol } H_2O}} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}} = 132 \text{ g } CO_2$$

4 *EVALUATE* The answer is rounded correctly to three significant figures to match those in 3.00 mol H_2O. The units cancel to leave g CO_2, which is the unknown. The answer is close to an estimate of 120, which is 3×40.

PRACTICE *Answers in Appendix E*

1. When magnesium burns in air, it combines with oxygen to form magnesium oxide according to the following equation.

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

What mass in grams of magnesium oxide is produced from 2.00 mol of magnesium?

2. What mass of glucose can be produced from a photosynthesis reaction that occurs using 10 mol CO_2?

$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate unknown quantities by using mole ratios.

Keyword: HC6STCX

Conversions of Mass to Amounts in Moles

In these stoichiometric calculations, you are asked to calculate the amount in moles of one substance that will react with or be produced from a given mass of another substance. In this type of problem, you are starting with a mass (probably in grams) of some substance. The plan for this conversion is

$$\begin{matrix} \text{mass of} & & \text{amount of} & & \text{amount of} \\ \textit{given}\text{ substance} & \longrightarrow & \textit{given}\text{ substance} & \longrightarrow & \textit{unknown}\text{ substance} \\ \text{(g)} & & \text{(mol)} & & \text{(mol)} \end{matrix}$$

This route requires two additional pieces of data: the molar mass of the *given* substance and the mole ratio. The molar mass is determined by using masses from the periodic table. We will follow a procedure much like the one used previously by using the units of the molar mass conversion factor to guide our mathematical operations. Because the known quantity is a mass, the conversion factor will need to be 1 mol divided by molar mass. This conversion factor cancels units of grams and leaves units of moles.

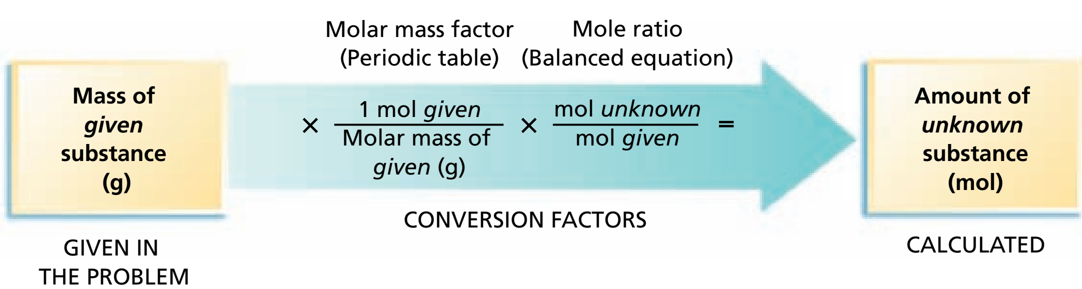


FIGURE 3 This is a solution plan for problems in which the given quantity is expressed in grams and the unknown quantity is expressed in moles.

SAMPLE PROBLEM D

For more help, go to the Math Tutor at the end of this chapter.

The first step in the industrial manufacture of nitric acid is the catalytic oxidation of ammonia.

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g) \quad \text{(unbalanced)}$$

The reaction is run using 824 g NH_3 and excess oxygen.
a. How many moles of NO are formed?
b. How many moles of H_2O are formed?

SOLUTION

1 ANALYZE

Given: mass of NH_3 = 824 g

Unknown: **a.** amount of NO produced (mol)
b. amount of H_2O produced (mol)

2 PLAN

First, write the balanced chemical equation.

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

Two conversion factors are needed to solve part (a)—the molar mass factor for NH_3 and the mole ratio of NO to NH_3. Part (b) starts with the same conversion factor as part (a), but then the mole ratio of H_2O to NH_3 is used to convert to the amount in moles of H_2O. The first conversion factor in each part is the molar mass factor of NH_3.

$$\textbf{a. } \text{g } NH_3 \times \overset{\text{molar mass factor}}{\frac{1 \text{ mol } NH_3}{\text{g } NH_3}} \times \overset{\text{mol ratio}}{\frac{\text{mol NO}}{\text{mol } NH_3}} = \text{mol NO}$$

$$\textbf{b. } \text{g } NH_3 \times \overset{\text{molar mass factor}}{\frac{1 \text{ mol } NH_3}{\text{g } NH_3}} \times \overset{\text{mol ratio}}{\frac{\text{mol } H_2O}{\text{mol } NH_3}} = \text{mol } H_2O$$

3 COMPUTE

Use the periodic table to compute the molar mass of NH_3.

$$1 \text{ mol } NH_3 = 17.04 \text{ g/mol}$$

$$\textbf{a. } 824 \cancel{\text{g } NH_3} \times \frac{1 \cancel{\text{mol } NH_3}}{17.04 \cancel{\text{g } NH_3}} \times \frac{4 \text{ mol NO}}{4 \cancel{\text{mol } NH_3}} = 48.4 \text{ mol NO}$$

$$\textbf{b. } 824 \cancel{\text{g } NH_3} \times \frac{1 \cancel{\text{mol } NH_3}}{17.04 \cancel{\text{g } NH_3}} \times \frac{6 \text{ mol } H_2O}{4 \cancel{\text{mol } NH_3}} = 72.5 \text{ mol } H_2O$$

4 EVALUATE

The answers are correctly given to three significant figures. The units cancel in the two problems to leave mol NO and mol H_2O, respectively, which are the unknowns.

PRACTICE

Answers in Appendix E

Oxygen was discovered by Joseph Priestley in 1774 when he heated mercury(II) oxide to decompose it to form its constituent elements.

1. How many moles of mercury(II) oxide, HgO, are needed to produce 125 g of oxygen, O_2?
2. How many moles of mercury are produced?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate unknown quantities by using mole ratios.

Keyword: HC6STCX

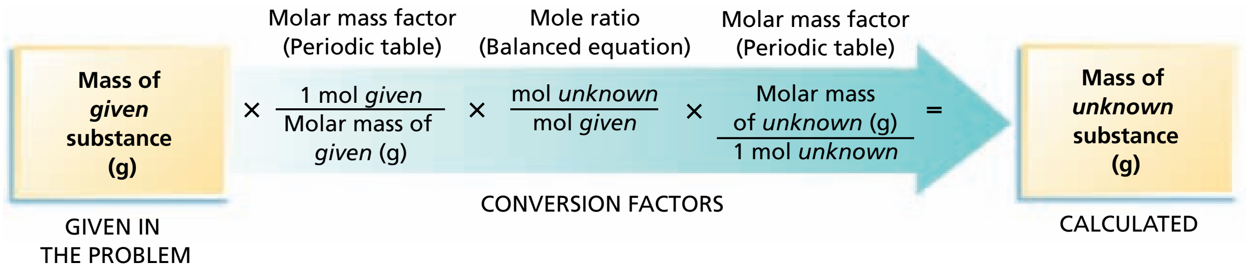


FIGURE 4 This is a solution plan for problems in which the given quantity is expressed in grams and the unknown quantity is also expressed in grams.

Mass-Mass Calculations

Mass-mass calculations are more practical than other mole calculations you have studied. You can never measure moles directly. You are generally required to calculate the amount in moles of a substance from its mass, which you can measure in the lab. Mass-mass problems can be viewed as the combination of the other types of problems. The plan for solving mass-mass problems is

$$\begin{array}{c}\text{mass of}\\ \textit{given}\text{ substance}\\ \text{(g)}\end{array} \longrightarrow \begin{array}{c}\text{amount of}\\ \textit{given}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{amount of}\\ \textit{unknown}\text{ substance}\\ \text{(mol)}\end{array} \longrightarrow \begin{array}{c}\text{mass of}\\ \textit{unknown}\text{ substance}\\ \text{(g)}\end{array}$$

Three additional pieces of data are needed to solve mass-mass problems: the molar mass of the *given* substance, the mole ratio, and the molar mass of the *unknown* substance.

SAMPLE PROBLEM E *For more help, go to the Math Tutor at the end of this chapter.*

Tin(II) fluoride, SnF_2, is used in some toothpastes. It is made by the reaction of tin with hydrogen fluoride according to the following equation.

$$\mathbf{Sn}(s) + \mathbf{2HF}(g) \longrightarrow \mathbf{SnF_2}(s) + \mathbf{H_2}(g)$$

How many grams of SnF_2 are produced from the reaction of 30.00 g HF with Sn?

SOLUTION

1 *ANALYZE*

Given: amount of HF = 30.00 g
Unknown: mass of SnF_2 produced (g)

2 *PLAN*

The conversion factors needed are the molar masses of HF and SnF_2 and the mole ratio of SnF_2 to HF.

$$\text{g HF} \times \overset{\text{molar mass factor}}{\frac{\text{mol HF}}{\text{g HF}}} \times \overset{\text{mol ratio}}{\frac{\text{mol SnF}_2}{\text{mol HF}}} \times \overset{\text{molar mass factor}}{\frac{\text{g SnF}_2}{\text{mol SnF}_2}} = \text{g SnF}_2$$

3 COMPUTE Use the periodic table to compute the molar masses of HF and SnF_2.

$$1 \text{ mol HF} = 20.01 \text{ g}$$
$$1 \text{ mol } SnF_2 = 156.71 \text{ g}$$

$$30.00 \text{ g } \cancel{\text{HF}} \times \frac{1 \cancel{\text{ mol HF}}}{20.01 \text{ g } \cancel{\text{HF}}} \times \frac{1 \cancel{\text{ mol } SnF_2}}{2 \cancel{\text{ mol HF}}} \times \frac{156.71 \text{ g } SnF_2}{1 \cancel{\text{ mol } SnF_2}} = 117.5 \text{ g } SnF_2$$

4 EVALUATE The answer is correctly rounded to four significant figures. The units cancel to leave g SnF_2, which matches the unknown. The answer is close to an estimated value of 120.

PRACTICE *Answers in Appendix E*

1. Laughing gas (nitrous oxide, N_2O) is sometimes used as an anesthetic in dentistry. It is produced when ammonium nitrate is decomposed according to the following reaction.

$$NH_4NO_3(s) \longrightarrow N_2O(g) + 2H_2O(l)$$

   a. How many grams of NH_4NO_3 are required to produce 33.0 g N_2O?

   b. How many grams of water are produced in this reaction?

2. When copper metal is added to silver nitrate in solution, silver metal and copper(II) nitrate are produced. What mass of silver is produced from 100. g Cu?

3. What mass of aluminum is produced by the decomposition of 5.0 kg Al_2O_3?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate unknown quantities by using mole ratios.

Keyword: HC6STCX

SECTION REVIEW

1. Balance the following equation. Then, given the moles of reactant or product below, determine the corresponding amount in moles of each of the other reactants and products.

$$NH_3 + O_2 \longrightarrow N_2 + H_2O$$

   a. 4 mol NH_3 **b.** 4 mol N_2 **c.** 4.5 mol O_2

2. One reaction that produces hydrogen gas can be represented by the following unbalanced chemical equation:

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

   a. What mass of HCl is consumed by the reaction of 2.50 moles of magnesium?

   b. What mass of each product is produced in part (a)?

3. Acetylene gas, C_2H_2, is produced as a result of the following reaction:

$$CaC_2(s) + 2H_2O(l) \longrightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

   a. If 32.0 g CaC_2 are consumed in this reaction, how many moles of H_2O are needed?

   b. How many moles of each product would form?

4. When sodium chloride reacts with silver nitrate, silver chloride precipitates. What mass of AgCl is produced from 75.0 g $AgNO_3$?

Critical Thinking

5. **RELATING IDEAS** Carbon and oxygen react to form carbon monoxide: $2C + O_2 \longrightarrow 2CO$. What masses of carbon and oxygen are needed to make 56.0 g CO? Which law does this illustrate?

SECTION 3

Limiting Reactants and Percentage Yield

OBJECTIVES

- Describe a method for determining which of two reactants is a limiting reactant.
- Calculate the amount in moles or mass in grams of a product, given the amounts in moles or masses in grams of two reactants, one of which is in excess.
- Distinguish between theoretical yield, actual yield, and percentage yield.
- Calculate percentage yield, given the actual yield and quantity of a reactant.

In the laboratory, a reaction is rarely carried out with exactly the required amount of each of the reactants. In many cases, one or more reactants is present in excess; that is, there is more than the exact amount required to react.

Once one of the reactants is used up, no more product can be formed. The substance that is completely used up first in a reaction is called the limiting reactant. *The* **limiting reactant** *is the reactant that limits the amount of the other reactant that can combine and the amount of product that can form in a chemical reaction. The substance that is not used up completely in a reaction is called the* **excess reactant.** A limiting reactant may also be referred to as a *limiting reagent.*

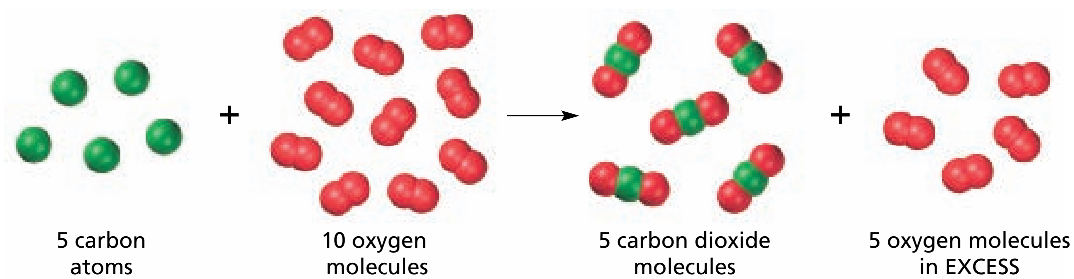
The concept of the limiting reactant is analogous to the relationship between the number of bicycles that can be made from the number of frames and wheels available. How many bicycles can you make if you have 100 frames and 250 wheels? You have frames to make 100 bicycles and wheels to make 125. The maximum number is *limited* by the number of frames, 100. Because you put two wheels on each frame, you will use 200 wheels and have 50 wheels left over.

The same reasoning can be applied to chemical reactions. Consider the reaction between carbon and oxygen to form carbon dioxide.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

According to the equation, one mole of carbon reacts with one mole of oxygen to form one mole of carbon dioxide. Suppose you could mix 5 mol C with 10 mol O_2 and allow the reaction to take place. **Figure 5** shows that there is more oxygen than is needed to react with the carbon. Carbon is the limiting reactant in this situation, and it limits the amount of CO_2 that is formed. Oxygen is the excess reactant, and 5 mol O_2 will be left over at the end of the reaction.

FIGURE 5 If you think of a mole as a multiple of molecules and atoms, you can see why the amount of O_2 is in excess.



SAMPLE PROBLEM F

Silicon dioxide (quartz) is usually quite unreactive but reacts readily with hydrogen fluoride according to the following equation.

$$SiO_2(s) + 4HF(g) \longrightarrow SiF_4(g) + 2H_2O(l)$$

If 6.0 mol HF is added to 4.5 mol SiO_2, which is the limiting reactant?

SOLUTION

1 ANALYZE

Given: amount of HF = 6.0 mol
amount of SiO_2 = 4.5 mol

Unknown: limiting reactant

2 PLAN

Pick one of the products, in this case SiF_4. Use the given amounts of each reactant to calculate the amount of SiF_4 that could be produced from that reactant. Compare the amounts of SiF_4. The limiting reactant is the reactant that produces the smallest number of moles of SiF_4. The smallest amount of product is also the maximum amount that can be formed.

$$\text{mol HF} \times \frac{\text{mol SiF}_4}{\text{mol HF}} = \text{mol SiF}_4 \text{ produced} \qquad \text{mol SiO}_2 \times \frac{\text{mol SiF}_4}{\text{mol SiO}_2} = \text{mol SiF}_4 \text{ produced}$$

3 COMPUTE

$$6.0 \cancel{\text{ mol HF}} \times \frac{1 \text{ mol SiF}_4}{4 \cancel{\text{ mol HF}}} = 1.5 \text{ mol SiF}_4 \text{ produced}$$

$$4.5 \cancel{\text{ mol SiO}_2} \times \frac{1 \text{ mol SiF}_4}{1 \cancel{\text{ mol SiO}_2}} = 4.5 \text{ mol SiF}_4 \text{ produced}$$

Under ideal conditions, 6.0 mol HF can make 1.5 mol SiF_4, and 4.5 mol SiO_2 present can make 4.5 mol SiF_4. Because 1.5 mol SiF_4 is smaller than 4.5 mol SiF_4, the HF is the limiting reactant and SiO_2 is the excess reactant.

4 EVALUATE

From the balanced equation, we can see that the reaction requires four times the number of moles of HF as it does moles of SiO_2. Because the molar amount of HF that we have is less than four times the moles of SiO_2, our calculations clearly show that HF is the limiting reactant.

PRACTICE

Answers in Appendix E

1. Some rocket engines use a mixture of hydrazine, N_2H_4, and hydrogen peroxide, H_2O_2, as the propellant. The reaction is given by the following equation.

$$N_2H_4(l) + 2H_2O_2(l) \longrightarrow N_2(g) + 4H_2O(g)$$

a. Which is the limiting reactant in this reaction when 0.750 mol N_2H_4 is mixed with 0.500 mol H_2O_2?

b. How much of the excess reactant, in moles, remains unchanged?

c. How much of each product, in moles, is formed?

extension

Go to **go.hrw.com** for more practice problems that ask you to determine the limiting reactant.

Keyword: HC6STCX

SAMPLE PROBLEM G

The black oxide of iron, Fe_3O_4, occurs in nature as the mineral magnetite. This substance can also be made in the laboratory by the reaction between red-hot iron and steam according to the following equation.

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

a. When 36.0 g H_2O is mixed with 67.0 g Fe, which is the limiting reactant?
b. What mass in grams of black iron oxide is produced?
c. What mass in grams of excess reactant remains when the reaction is completed?

SOLUTION

1 ANALYZE

Given: mass of H_2O = 36.0 g
mass of Fe = 67.0 g

Unknown: limiting reactant
mass of Fe_3O_4, in grams
mass of excess reactant remaining

2 PLAN

a. First, convert both given masses in grams to amounts in moles. Then, calculate the number of moles of one of the products. Because the problem asks for the mass of Fe_3O_4 formed, we will calculate moles of Fe_3O_4. The reactant yielding the smaller number of moles of product is the limiting reactant.

$$\text{g Fe} \times \overset{\text{molar mass factor}}{\frac{\text{mol Fe}}{\text{g Fe}}} \times \overset{\text{mol ratio}}{\frac{\text{mol } Fe_3O_4}{\text{mol Fe}}} = \text{mol } Fe_3O_4$$

$$\text{g } H_2O \times \overset{\text{molar mass factor}}{\frac{\text{mol } H_2O}{\text{g } H_2O}} \times \overset{\text{mol ratio}}{\frac{\text{mol } Fe_3O_4}{\text{mol } H_2O}} = \text{mol } Fe_3O_4$$

b. To find the maximum mass of Fe_3O_4 that can be produced, we must use the amount of Fe_3O_4 in moles from the limiting reactant in a simple stoichiometric problem.

$$\text{mole } Fe_3O_4 \text{ from limiting reactant} \times \overset{\text{molar mass factor}}{\frac{\text{g } Fe_3O_4}{\text{mol } Fe_3O_4}} = \text{g } Fe_3O_4 \text{ produced}$$

c. To find the amount of excess reactant remaining, we must first determine the amount of the excess reactant that is consumed. The calculated moles of the product (from the limiting reactant) is used to determine the amount of excess reactant that is consumed.

$$\text{mol product} \times \frac{\text{mol excess reactant}}{\text{mol product}} \times \frac{\text{g excess reactant}}{\text{mol excess reactant}} = \text{g excess reactant consumed}$$

The amount of excess reactant remaining can then be found by subtracting the amount consumed from the amount originally present.

original g excess reactant – g excess reactant consumed = g excess reactant remaining

3 *COMPUTE*

a. Use the periodic table to determine the molar masses of H_2O, Fe, and Fe_3O_4. Then, determine how many mol Fe_3O_4 can be produced from each reactant.

1 mol H_2O = 18.02 g

1 mol Fe = 55.85 g

1 mol Fe_3O_4 = 231.55 g

$$67.0\ \cancel{\text{g Fe}} \times \frac{1\ \cancel{\text{mol Fe}}}{55.85\ \cancel{\text{g Fe}}} \times \frac{1\ \text{mol Fe}_3\text{O}_4}{3\ \cancel{\text{mol Fe}}} = 0.400\ \text{mol Fe}_3\text{O}_4$$

$$36.0\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \cancel{\text{mol H}_2\text{O}}}{18.02\ \cancel{\text{g H}_2\text{O}}} \times \frac{1\ \text{mol Fe}_3\text{O}_4}{4\ \cancel{\text{mol H}_2\text{O}}} = 0.499\ \text{mol Fe}_3\text{O}_4$$

Fe is the limiting reactant because the given amount of Fe can make only 0.400 mol Fe_3O_4, which is less than the 0.499 mol Fe_3O_4 that the given amount of H_2O would produce.

b. $$0.400\ \text{mol Fe}_3\text{O}_4 \times \frac{231.55\ \text{g Fe}_3\text{O}_4}{1\ \text{mol Fe}_3\text{O}_4} = 92.6\ \text{g Fe}_3\text{O}_4$$

c. $$0.400\ \text{mol Fe}_3\text{O}_4 \times \frac{4\ \text{mol H}_2\text{O}}{1\ \text{mol Fe}_3\text{O}_4} \times \frac{18.02\ \text{g H}_2\text{O}}{1\ \text{mol H}_2\text{O}} = 28.8\ \text{g H}_2\text{O consumed}$$

36.0 g H_2O – 28.8 g H_2O consumed = 7.2 g H_2O remaining

4 *EVALUATE*

The mass of original reactants is 67.0 + 36.0 = 103.0 g; the mass of Fe_3O_4 + unreacted water is 92.6 g + 7.2 g = 99.8 g. The difference of 3.2 g is the mass of hydrogen that is produced with the Fe_3O_4.

PRACTICE

Answers in Appendix E

1. Zinc and sulfur react to form zinc sulfide according to the following equation.

$$8\text{Zn}(s) + \text{S}_8(s) \longrightarrow 8\text{ZnS}(s)$$

a. If 2.00 mol of Zn are heated with 1.00 mol of S_8, identify the limiting reactant.

b. How many moles of excess reactant remain?

c. How many moles of the product are formed?

2. Carbon reacts with steam, H_2O, at high temperatures to produce hydrogen and carbon monoxide.

a. If 2.40 mol of carbon are exposed to 3.10 mol of steam, identify the limiting reactant.

b. How many moles of each product are formed?

c. What mass of each product is formed?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate the amount of excess reactant and the amount of product formed.

Keyword: HC6STCX

QuickLAB

Wear oven mitts when handling heated items.

Limiting Reactants in a Recipe

Materials

- 1/2 cup sugar
- 1/2 cup brown sugar
- 1 1/3 stick margarine (at room temperature)
- 1 egg
- 1/2 tsp. salt
- 1 tsp. vanilla
- 1/2 tsp. baking soda
- 1 1/2 cup flour
- 1 1/3 cup chocolate chips
- mixing bowl
- mixing spoon
- measuring spoons and cups
- cookie sheet
- oven preheated to 350°F

Procedure

1. In the mixing bowl, combine the sugars and margarine together until smooth. (An electric mixer will make this process go much faster.)
2. Add the egg, salt, and vanilla. Mix well.
3. Stir in the baking soda, flour, and chocolate chips. Chill the dough for an hour in the refrigerator for best results.
4. Divide the dough into 24 small balls about 3 cm in diameter. Place the balls on an ungreased cookie sheet.
5. Bake at 350°F for about 10 minutes, or until the cookies are light brown.

Yield: 24 cookies

Discussion

1. Suppose you are given the following amounts of ingredients:
   1 dozen eggs
   24 tsp. of vanilla
   1 lb. (82 tsp.) of salt
   1 lb. (84 tsp.) of baking soda
   3 cups of chocolate chips
   5 lb. (11 cups) of sugar
   2 lb. (4 cups) of brown sugar
   1 lb. (4 sticks) of margarine

   a. For each ingredient, calculate how many cookies could be prepared if all of that ingredient were consumed. (For example, the recipe shows that using 1 egg—with the right amounts of the other ingredients—yields 24 cookies. How many cookies can you make if the recipe is increased proportionately for 12 eggs?)

   b. To determine the limiting reactant for the new ingredients list, identify which ingredient will result in the fewest number of cookies.

   c. What is the maximum number of cookies that can be produced from the new amounts of ingredients?

Percentage Yield

The amounts of products calculated in the ideal stoichiometry problems in this chapter so far represent theoretical yields. *The* **theoretical yield** *is the maximum amount of product that can be produced from a given amount of reactant.* In most chemical reactions, the amount of product obtained is less than the theoretical yield. There are many reasons for this result. Reactants may contain impurities or may form byproducts in competing side reactions. Also, in many reactions, all reactants are not converted to products. As a result, less product is produced than ideal stoichiometric calculations predict. *The measured amount of a product obtained from a reaction is called the* **actual yield** *of that product.*

Chemists are usually interested in the efficiency of a reaction. The efficiency is expressed by comparing the actual and theoretical yields. *The* **percentage yield** *is the ratio of the actual yield to the theoretical yield, multiplied by 100.*

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

SAMPLE PROBLEM H

Chlorobenzene, C_6H_5Cl, is used in the production of many important chemicals, such as aspirin, dyes, and disinfectants. One industrial method of preparing chlorobenzene is to react benzene, C_6H_6, with chlorine, as represented by the following equation.

$$C_6H_6(l) + Cl_2(g) \longrightarrow C_6H_5Cl(l) + HCl(g)$$

When 36.8 g C_6H_6 react with an excess of Cl_2, the actual yield of C_6H_5Cl is 38.8 g. What is the percentage yield of C_6H_5Cl?

SOLUTION

1 ANALYZE

Given: mass of C_6H_6 = 36.8 g
mass of Cl_2 = excess
actual yield of C_6H_5Cl = 38.8 g
Unknown: percentage yield of C_6H_5Cl

2 PLAN

First do a mass-mass calculation to find the theoretical yield of C_6H_5Cl.

$$\text{g } C_6H_6 \times \underbrace{\frac{\text{mol } C_6H_6}{\text{g } C_6H_6}}_{\text{molar mass factor}} \times \underbrace{\frac{\text{mol } C_6H_5Cl}{\text{mol } C_6H_6}}_{\text{mol ratio}} \times \underbrace{\frac{\text{g } C_6H_5Cl}{\text{mol } C_6H_5Cl}}_{\text{molar mass}} = \text{g } C_6H_5Cl \text{ (theoretical yield)}$$

Then the percentage yield can be found.

$$\text{percentage yield } C_6H_5Cl = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

3 COMPUTE Use the periodic table to determine the molar masses of C_6H_6 and C_6H_5Cl.

$$1 \text{ mol } C_6H_6 = 78.12 \text{ g}$$
$$1 \text{ mol } C_6H_5Cl = 112.56 \text{ g}$$

$$36.8 \text{ g } \cancel{C_6H_6} \times \frac{1 \text{ mol } \cancel{C_6H_6}}{78.12 \text{ g } \cancel{C_6H_6}} \times \frac{1 \text{ mol } \cancel{C_6H_5Cl}}{1 \text{ mol } \cancel{C_6H_6}} \times \frac{112.56 \text{ g } C_6H_5Cl}{1 \text{ mol } \cancel{C_6H_5Cl}} = 53.0 \text{ g } C_6H_5Cl \text{ (theoretical yield)}$$

$$\text{percentage yield} = \frac{38.8 \text{ g}}{53.0 \text{ g}} \times 100 = 73.2\%$$

4 EVALUATE The answer is correctly rounded to three significant figures to match those in 36.8 g C_6H_6. The units have canceled correctly. The theoretical yield is close to an estimated value of 50 g, (one-half of 100 g). The percentage yield is close to an estimated value of 80%, (40/50 × 100).

PRACTICE *Answers in Appendix E*

1. Methanol can be produced through the reaction of CO and H_2 in the presence of a catalyst.

$$CO(g) + 2H_2(g) \xrightarrow{\text{catalyst}} CH_3OH(l)$$

If 75.0 g of CO reacts to produce 68.4 g CH_3OH, what is the percentage yield of CH_3OH?

2. Aluminum reacts with excess copper(II) sulfate according to the reaction given below. If 1.85 g of Al react and the percentage yield of Cu is 56.6%, what mass of Cu is produced?

$$Al(s) + CuSO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + Cu(s) \text{ (unbalanced)}$$

extension
Go to **go.hrw.com** for more practice problems that ask you to calculate percentage yield.
Keyword: HC6STCX

SECTION REVIEW

1. Carbon disulfide burns in oxygen to yield carbon dioxide and sulfur dioxide according to the following chemical equation.

$$CS_2(l) + 3O_2(g) \longrightarrow CO_2(g) + 2SO_2(g)$$

   a. If 1.00 mol CS_2 reacts with 1.00 mol O_2, identify the limiting reactant.

   b. How many moles of excess reactant remain?

   c. How many moles of each product are formed?

2. Metallic magnesium reacts with steam to produce magnesium hydroxide and hydrogen gas.

   a. If 16.2 g Mg are heated with 12.0 g H_2O, what is the limiting reactant?

   b. How many moles of the excess reactant are left?

   c. How many grams of each product are formed?

3. Quicklime, CaO, can be prepared by roasting limestone, $CaCO_3$, according to the following reaction.

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

When 2.00×10^3 g $CaCO_3$ are heated, the actual yield of CaO is 1.05×10^3 g. What is the percentage yield?

Critical Thinking

4. **ANALYZING DATA** A chemical engineer calculated that 15.0 mol H_2 was needed to react with excess N_2 to prepare 10.0 mol NH_3. But the actual yield is 60.0%. Write a balanced chemical equation for the reaction. Is the amount of H_2 needed to make 10.0 mol NH_3 more, the same, or less than 15 mol? How many moles of H_2 are needed?

CHAPTER HIGHLIGHTS

Introduction to Stoichiometry

Vocabulary
composition stoichiometry
reaction stoichiometry
mole ratio

- Reaction stoichiometry involves the mass relationships between reactants and products in a chemical reaction.
- Relating one substance to another requires expressing the amount of each substance in moles.
- A mole ratio is the conversion factor that relates the amount in moles of any two substances in a chemical reaction. The mole ratio is derived from the balanced equation.
- Amount of a substance is expressed in moles, and mass of a substance is expressed by using mass units such as grams, kilograms, or milligrams.
- Mass and amount of substance are quantities, whereas moles and grams are units.
- A balanced chemical equation is necessary to solve any stoichiometric problem.

Ideal Stoichiometric Calculations

- In an ideal stoichiometric calculation, the mass or the amount of any reactant or product can be calculated if the balanced chemical equation and the mass or amount of any other reactant or product is known.

Limiting Reactants and Percentage Yield

Vocabulary
limiting reactant
excess reactant
theoretical yield
actual yield
percentage yield

- In actual reactions, the reactants may be present in proportions that differ from the stoichiometric proportions required for a complete reaction in which all of each reactant is converted to product.
- The limiting reactant controls the maximum possible amount of product formed.
- For many reactions, the quantity of a product is less than the theoretical maximum for that product. Percentage yield shows the relationship between the theoretical yield and actual yield for the product of a reaction.

CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

Introduction to Stoichiometry

SECTION 1 REVIEW

1. a. Explain the concept of mole ratio as used in reaction stoichiometry problems.
   b. What is the source of this ratio?
2. For each of the following balanced chemical equations, write all possible mole ratios:
   a. $2Ca + O_2 \longrightarrow 2CaO$
   b. $Mg + 2HF \longrightarrow MgF_2 + H_2$

PRACTICE PROBLEMS

3. Given the chemical equation $Na_2CO_3(aq) + Ca(OH)_2 \longrightarrow 2NaOH(aq) + CaCO_3(s)$, determine to two decimal places the molar masses of all substances involved. Then, write the molar masses as conversion factors.

Ideal Stoichiometric Calculations

SECTION 2 REVIEW

4. a. What is molar mass?
   b. What is its role in reaction stoichiometry?

PRACTICE PROBLEMS

5. Hydrogen and oxygen react under a specific set of conditions to produce water according to the following: $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$.
   a. How many moles of hydrogen would be required to produce 5.0 mol of water?
   b. How many moles of oxygen would be required? (Hint: See Sample Problem A.)
6. a. If 4.50 mol of ethane, C_2H_6, undergo combustion according to the unbalanced equation $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$, how many moles of oxygen are required?
   b. How many moles of each product are formed?
7. Sodium chloride is produced from its elements through a synthesis reaction. What mass of each reactant would be required to produce 25.0 mol of sodium chloride?
8. In a blast furnace, iron(lll) oxide is used to produce iron by the following (unbalanced) reaction: $Fe_2O_3(s) + CO(g) \longrightarrow Fe(s) + CO_2(g)$
   a. If 4.00 kg Fe_2O_3 are available to react, how many moles of CO are needed?
   b. How many moles of each product are formed?
9. Methanol, CH_3OH, is an important industrial compound that is produced from the following (unbalanced) reaction: $CO(g) + H_2(g) \longrightarrow CH_3OH(g)$. What mass of each reactant would be needed to produce 100.0 kg of methanol? (Hint: See Sample Problem E.)
10. Nitrogen combines with oxygen in the atmosphere during lightning flashes to form nitrogen monoxide, NO, which then reacts further with O_2 to produce nitrogen dioxide, NO_2.
    a. What mass of NO_2 is formed when NO reacts with 384 g O_2?
    b. How many grams of NO are required to react with this amount of O_2?
11. As early as 1938, the use of NaOH was suggested as a means of removing CO_2 from the cabin of a spacecraft according to the following (unbalanced) reaction: $NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$.
    a. If the average human body discharges 925.0 g CO_2 per day, how many moles of NaOH are needed each day for each person in the spacecraft?
    b. How many moles of each product are formed?
12. The double-replacement reaction between silver nitrate and sodium bromide produces silver bromide, a component of photographic film.
    a. If 4.50 mol of silver nitrate react, what mass of sodium bromide is required?
    b. What mass of silver bromide is formed?
13. In a soda-acid fire extinguisher, concentrated sulfuric acid reacts with sodium hydrogen carbonate to produce carbon dioxide, sodium sulfate, and water.
    a. How many moles of sodium hydrogen carbonate would be needed to react with 150.0 g of sulfuric acid?
    b. How many moles of each product would be formed?

14. Sulfuric acid reacts with sodium hydroxide according to the following:
$H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O$.
a. Balance the equation for this reaction.
b. What mass of H_2SO_4 would be required to react with 0.75 mol NaOH?
c. What mass of each product is formed by this reaction? (Hint: See Sample B.)

15. Copper reacts with silver nitrate through single replacement.
a. If 2.25 g of silver are produced from the reaction, how many moles of copper(II) nitrate are also produced?
b. How many moles of each reactant are required in this reaction? (Hint: See Sample Problem D.)

16. Aspirin, $C_9H_8O_4$, is produced through the following reaction of salicylic acid, $C_7H_6O_3$, and acetic anhydride, $C_4H_6O_3$: $C_7H_6O_3(s) + C_4H_6O_3(l) \longrightarrow C_9H_8O_4(s) + HC_2H_3O_2(l)$.
a. What mass of aspirin (kg) could be produced from 75.0 mol of salicylic acid?
b. What mass of acetic anhydride (kg) would be required?
c. At 20°C, how many liters of acetic acid, $HC_2H_3O_2$, would be formed? The density of $HC_2H_3O_2$ is 1.05 g/mL.

Limiting Reactants and Percentage Yield

SECTION 3 REVIEW

17. Distinguish between ideal and real stoichiometric calculations.

18. Distinguish between the limiting reactant and the excess reactant in a chemical reaction.

19. a. Distinguish between the theoretical yield and actual yield in stoichiometric calculations.
b. How does the value of the theoretical yield generally compare with the value of the actual yield?

20. What is the percentage yield of a reaction?

21. Why are actual yields usually less than calculated theoretical yields?

PRACTICE PROBLEMS

22. Given the reactant amounts specified in each chemical equation, determine the limiting reactant in each case:
a. $\underset{2.0\ mol}{HCl} + \underset{2.5\ mol}{NaOH} \longrightarrow NaCl + H_2O$
b. $\underset{2.5\ mol}{Zn} + \underset{6.0\ mol}{2HCl} \longrightarrow ZnCl_2 + H_2$
c. $\underset{4.0\ mol}{2Fe(OH)_3} + \underset{6.5\ mol}{3H_2SO_4} \longrightarrow Fe_2(SO_4)_3 + 6H_2O$
(Hint: See Sample Problem F.)

23. For each reaction specified in Problem 22, determine the amount in moles of excess reactant that remains. (Hint: See Sample Problem G.)

24. For each reaction specified in Problem 22, calculate the amount in moles of each product formed.

25. a. If 2.50 mol of copper and 5.50 mol of silver nitrate are available to react by single replacement, identify the limiting reactant.
b. Determine the amount in moles of excess reactant remaining.
c. Determine the amount in moles of each product formed.
d. Determine the mass of each product formed.

26. Sulfuric acid reacts with aluminum hydroxide by double replacement.
a. If 30.0 g of sulfuric acid react with 25.0 g of aluminum hydroxide, identify the limiting reactant.
b. Determine the mass of excess reactant remaining.
c. Determine the mass of each product formed. Assume 100% yield.

27. The energy used to power one of the Apollo lunar missions was supplied by the following overall reaction: $2N_2H_4 + (CH_3)_2N_2H_2 + 3N_2O_4 \longrightarrow 6N_2 + 2CO_2 + 8H_2O$. For the phase of the mission when the lunar module ascended from the surface of the moon, a total of 1200. kg N_2H_4 was available to react with 1000. kg $(CH_3)_2N_2H_2$ and 4500. kg N_2O_4.
a. For this portion of the flight, which of the allocated components was used up first?
b. How much water, in kilograms, was put into the lunar atmosphere through this reaction?

28. Calculate the indicated quantity for each of the various chemical reactions given:

a. theoretical yield = 20.0 g, actual yield = 15.0 g, percentage yield = ?

b. theoretical yield = 1.0 g, percentage yield = 90.0%, actual yield = ?

c. theoretical yield = 5.00 g, actual yield = 4.75 g, percentage yield = ?

d. theoretical yield = 3.45 g, percentage yield = 48.0%, actual yield = ?

29. The percentage yield for the reaction

$$PCl_3 + Cl_2 \longrightarrow PCl_5$$

is 83.2%. What mass of PCl_5 is expected from the reaction of 73.7 g PCl_3 with excess chlorine?

30. The Ostwald process for producing nitric acid from ammonia consists of the following steps:

$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$

$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

$3NO_2(g) + H_2O(g) \longrightarrow 2HNO_3(aq) + NO(g)$

If the yield in each step is 94.0%, how many grams of nitric acid can be produced from 5.00 kg of ammonia?

MIXED REVIEW

31. Magnesium is obtained from sea water. $Ca(OH)_2$ is added to sea water to precipitate $Mg(OH)_2$. The precipitate is filtered and reacted with HCl to produce $MgCl_2$. The $MgCl_2$ is electrolyzed to produce Mg and Cl_2. If 185.0 g of magnesium are recovered from 1000. g $MgCl_2$, what is the percentage yield for this reaction?

32. Phosphate baking powder is a mixture of starch, sodium hydrogen carbonate, and calcium dihydrogen phosphate. When mixed with water, phosphate baking powder releases carbon dioxide gas, causing a dough or batter to bubble and rise.

$$2NaHCO_3(aq) + Ca(H_2PO_4)_2(aq) \longrightarrow Na_2HPO_4(aq) + CaHPO_4(aq) + 2CO_2(g) + 2H_2O(l)$$

If 0.750 L CO_2 is needed for a cake and each kilogram of baking powder contains 168 g of $NaHCO_3$, how many grams of baking powder must be used to generate this amount CO_2? The density of CO_2 at baking temperature is about 1.20 g/L.

33. Coal gasification is a process that converts coal into methane gas. If this reaction has a percentage yield of 85.0%, what mass of methane can be obtained from 1250 g of carbon?

$$2C(s) + 2H_2O(l) \longrightarrow CH_4(g) + CO_2(g)$$

34. If the percentage yield for the coal gasification process is increased to 95%, what mass of methane can be obtained from 2750 g of carbon?

35. Builders and dentists must store plaster of Paris, $CaSO_4 \cdot \frac{1}{2}H_2O$, in airtight containers to prevent it from absorbing water vapor from the air and changing to gypsum, $CaSO_4 \cdot 2H_2O$. How many liters of water vapor evolve when 2.00 kg of gypsum are heated at 110°C to produce plaster of Paris? At 110°C, the density of water vapor is 0.574 g/L.

36. Gold can be recovered from sea water by reacting the water with zinc, which is refined from zinc oxide. The zinc displaces the gold in the water. What mass of gold can be recovered if 2.00 g of ZnO and an excess of sea water are available?

$2ZnO(s) + C(s) \longrightarrow 2Zn(s) + CO_2(g)$

$2Au^{3+}(aq) + 3Zn(s) \longrightarrow 3Zn^{2+}(aq) + 2Au(s)$

CRITICAL THINKING

37. Relating Ideas The chemical equation is a good source of information concerning a reaction. Explain the relationship between the actual yield of a reaction product and the chemical equation of the product.

38. Analyzing Results Very seldom are chemists able to achieve a 100% yield of a product from a chemical reaction. However, the yield of a reaction is usually important because of the expense involved in producing less product. For example, when magnesium metal is heated in a crucible at high temperatures, the product magnesium oxide, MgO, is formed. Based on your analysis of the reaction, describe some of the actions that you would take to increase your percentage yield. The reaction is as follows:

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

39. **Analyzing Results** In the lab, you run an experiment that appears to have a percentage yield of 115%. Propose reasons for this result. Can an actual yield ever exceed a theoretical yield? Explain your answer.

40. **Relating Ideas** Explain the stoichiometry of blowing air on a smoldering campfire to keep the coals burning.

USING THE HANDBOOK

41. The steel-making process described in the Transition Metal section of the *Elements Handbook* shows the equation for the formation of iron carbide. Use this equation to answer the following questions:
    a. If 3.65×10^3 kg of iron is used in a steel-making process, what is the minimum mass of carbon needed to react with all of the iron?
    b. What is the theoretical mass of iron carbide that is formed?

42. The reaction of aluminum with oxygen to produce a protective coating for the metal's surface is described in the discussion of aluminum in Group 13 of the *Elements Handbook*. Use this equation to answer the following questions:
    a. What mass of aluminum oxide would theoretically be formed if a 30.0 g piece of aluminum foil reacted with excess oxygen?
    b. Why would you expect the actual yield from this reaction to be far less than the mass you calculated in item (a)?

43. The reactions of oxide compounds to produce carbonates, phosphates, and sulfates are described in the section on oxides in Group 16 of the *Elements Handbook*. Use those equations to answer the following questions:
    a. What mass of CO_2 is needed to react with 154.6 g MgO?
    b. What mass of magnesium carbonate is produced?
    c. When 45.7 g P_4O_{10} is reacted with an excess of calcium oxide, what mass of calcium phosphate is produced?

RESEARCH & WRITING

44. Research the history of the Haber process for the production of ammonia. What was the significance of this process in history? How is this process related to the discussion of reaction yields in this chapter?

ALTERNATIVE ASSESSMENT

45. **Performance** Just as reactants combine in certain proportions to form a product, colors can be combined to create other colors. Artists do this all the time to find just the right color for their paintings. Using poster paint, determine the proportions of primary pigments used to create the following colors. Your proportions should be such that anyone could mix the color perfectly.

46. **Performance** Write two of your own sample problems that are descriptions of how to solve a mass-mass problem. Assume that your sample problems will be used by other students to learn how to solve mass-mass problems.

extension

Graphing Calculator **Limiting Reactants and Percentage Yield**

Go to **go.hrw.com** for a graphing calculator exercise that asks you to use a theoretical yield graph to make predictions about limiting reactants and percentage yield.

Keyword: HC6STCX

Math Tutor USING MOLE RATIOS

An unbalanced chemical equation tells you what substances react and what products are produced. A balanced chemical equation gives you even more information. It tells you how many atoms, molecules, or ions react and how many atoms, molecules, or ions are produced. The coefficients in a balanced equation represent the relative amounts in moles of reactants and products. Using this information, you can set up a mole ratio. A mole ratio is a conversion factor that relates the amounts in moles of any two substances involved in a chemical reaction.

Problem-Solving TIPS

- When solving stoichiometric problems, always start with a balanced chemical equation.
- Identify the amount known from the problem (in moles or mass).
- If you are given the mass of a substance, use the molar mass factor as a conversion factor to find the amount in moles. If you are given the amount in moles of a substance, use the molar mass factor as a conversion factor to find the mass.

SAMPLE

If 3.61 g of aluminum reacts completely with excess $CuCl_2$, what mass of copper metal is produced? Use the balanced equation below.

$$2Al(s) + 3CuCl_2(aq) \longrightarrow 2AlCl_3(aq) + 3Cu(s)$$

You know the mass of aluminum that reacts. If you convert that mass to moles, you can apply the mole ratio of aluminum to copper in this reaction to find the moles of copper produced.

$$\text{mol Al} = 3.61\ \cancel{\text{g Al}} \times \frac{1\ \text{mol Al}}{26.98\ \cancel{\text{g Al}}} = 0.134\ \text{mol Al}$$

$$\cancel{\text{mol Al}} \times \frac{3\ \text{mol Cu}}{2\ \cancel{\text{mol Al}}} = \text{mol Cu}$$

$$0.134\ \cancel{\text{mol Al}} \times \frac{3\ \text{mol Cu}}{2\ \cancel{\text{mol Al}}} = 0.201\ \text{mol Cu}$$

Then, convert moles of Cu to mass of Cu by applying the following factor:

$$\text{mol Cu} \times \frac{\text{molar mass Cu}}{1\ \text{mol Cu}} = \text{mass Cu, or } 0.201\ \text{mol Cu} \times \frac{63.55\ \text{g Cu}}{1\ \text{mol Cu}} = 12.8\ \text{g Cu}$$

PRACTICE PROBLEMS

1. If 12.24 moles of O_2 react with excess SO_2, how many moles of SO_3 are formed? Use the balanced equation below.
$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

2. If 78.50 g $KClO_3$ decomposes, what mass of O_2 is produced? Use the balanced equation below.
$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. In stoichiometry, chemists are mainly concerned with

A. the types of bonds found in compounds.
B. mass relationships in chemical reactions.
C. energy changes occurring in chemical reactions.
D. the speed with which chemical reactions occur.

2. Assume ideal stoichiometry in the reaction $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$. If you know the mass of CH_4, you can calculate

A. only the mass of CO_2 produced.
B. only the mass of O_2 reacting.
C. only the mass of $CO_2 + H_2O$ produced.
D. the mass of O_2 reacting and $CO_2 + H_2O$ produced.

3. Which mole ratio for the equation $6Li + N_2 \longrightarrow 2Li_3N$ is incorrect?

A. $\frac{6 \text{ mol Li}}{2 \text{ mol } N_2}$ **C.** $\frac{2 \text{ mol } Li_3N}{1 \text{ mol } N_2}$
B. $\frac{1 \text{ mol } N_2}{6 \text{ mol Li}}$ **D.** $\frac{2 \text{ mol } Li_3N}{6 \text{ mol Li}}$

4. For the reaction below, how many moles of N_2 are required to produce 18 mol NH_3?

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

A. 4.5 **C.** 18
B. 9.0 **D.** 36

5. What mass of NaCl can be produced by the reaction of 0.75 mol Cl_2?

$$2Na + Cl_2 \longrightarrow 2NaCl$$

A. 0.75 g **C.** 44 g
B. 1.5 g **D.** 88 g

6. What mass of CO_2 can be produced from 25.0 g $CaCO_3$ given the decomposition reaction

$$CaCO_3 \longrightarrow CaO + CO_2$$

A. 11.0 g **C.** 25.0 g
B. 22.0 g **D.** 56.0 g

7. If a chemical reaction involving substances A and B stops when B is completely used up, then B is referred to as the

A. excess reactant. **C.** limiting reactant.
B. primary reactant. **D.** primary product.

8. If a chemist calculates the maximum amount of product that could be obtained in a chemical reaction, he or she is calculating the

A. percentage yield.
B. mole ratio.
C. theoretical yield.
D. actual yield.

9. What is the maximum number of moles of $AlCl_3$ that can be produced from 5.0 mol Al and 6.0 mol Cl_2?

$$2Al + 3Cl_2 \longrightarrow 2AlCl_3$$

A. 2.0 mol $AlCl_3$ **C.** 5.0 mol $AlCl_3$
B. 4.0 mol $AlCl_3$ **D.** 6.0 mol $AlCl_3$

SHORT ANSWER

10. Why is a balanced equation necessary to solve a mass-mass stoichiometry problem?

11. What data are necessary to calculate the percentage yield of a reaction?

EXTENDED RESPONSE

12. A student makes a compound in the laboratory and reports an actual yield of 120%. Is this result possible? Assuming that all masses were measured correctly, give an explanation.

13. Benzene, C_6H_6, is reacted with bromine, Br_2, to produce bromobenzene, C_6H_5Br, and hydrogen bromide, HBr, as shown below. When 40.0 g of benzene are reacted with 95.0 g of bromine, 65.0 g of bromobenzene is produced.

$$C_6H_6 + Br_2 \longrightarrow C_6H_5Br + HBr$$

a. Which compound is the limiting reactant?
b. What is the theoretical yield of bromobenzene?
c. What is the reactant in excess, and how much remains after the reaction is completed?
d. What is the percentage yield?

Test TIP Choose an answer to a question based on both information that you already know and information that is presented in the question.

CHAPTER LAB

SEE PRE-LAB
EXTRACTION AND FILTRATION

Stoichiometry and Gravimetric Analysis

OBJECTIVES

- ***Observe*** **the double-displacement reaction between solutions of strontium chloride and sodium carbonate.**
- ***Demonstrate*** **proficiency with gravimetric methods.**
- ***Measure*** **the mass of the precipitate that forms.**
- ***Relate*** **the mass of the precipitate that forms to the mass of the reactants before the reaction.**
- ***Calculate*** **the mass of sodium carbonate in a solution of unknown concentration.**

MATERIALS

- **15 mL Na_2CO_3 solution of unknown concentration**
- **50 mL 0.30 M $SrCl_2$ solution**
- **50 mL graduated cylinder**
- **250 mL beakers, 2**
- **balance**
- **beaker tongs**
- **distilled water**
- **drying oven**
- **filter paper**
- **glass funnel or Büchner funnel with related equipment**
- **glass stirring rod**
- **paper towels**
- **ring and ring stand**
- **spatula**
- **water bottle**

BACKGROUND

This gravimetric analysis involves a double-displacement reaction between strontium chloride, $SrCl_2$, and sodium carbonate, Na_2CO_3. This type of reaction can be used to determine the amount of a carbonate compound in a solution. For accurate results, essentially all of the reactant of unknown amount must be converted into product. If the mass of the product is carefully measured, you can use stoichiometric calculations to determine how much of the reactant of unknown amount was involved in the reaction.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Make a data table that has 2 columns and 5 rows. In the first column write each of the following phrases in a separate row: "Volume of Na_2CO_3 solution added"; "Volume of $SrCl_2$ solution added"; "Mass of dry filter paper"; "Mass of beaker with paper towel"; "Mass of beaker with paper towel, filter paper, and precipitate."
2. Clean all of the necessary lab equipment with soap and water, and rinse with distilled water.
3. Measure the mass of a piece of filter paper to the nearest 0.01 g, and record it in your table.
4. Set up a filtering apparatus. Use the Pre-Laboratory Procedure "Extraction and Filtration."

5. Label a paper towel with your name and the date. Place the towel in a clean, dry 250 mL beaker, and measure and record the mass of the paper towel and beaker to the nearest 0.01 g.

PROCEDURE

1. Measure about 15 mL of the Na_2CO_3 solution into the graduated cylinder. Record this volume to the nearest 0.5 mL. Pour the Na_2CO_3 solution into an empty 250 mL beaker. Carefully wash the graduated cylinder, and rinse it with distilled water.
2. Measure about 25 mL of the 0.30 M $SrCl_2$ solution into the graduated cylinder. Record this volume to the nearest 0.5 mL. Pour the $SrCl_2$ solution into the beaker with the Na_2CO_3 solution, as shown in **Figure A.** Gently stir with a glass stirring rod.
3. Measure another 10 mL of the $SrCl_2$ solution into the graduated cylinder. Record the volume to the nearest 0.5 mL. Slowly add the solution to the beaker, and stir gently. Repeat this step until no more precipitate forms.
4. Slowly pour the mixture into the funnel. Do not overfill the funnel—some of the precipitate could be lost between the filter paper and the funnel.
5. Rinse the beaker several more times with distilled water. Pour the rinse water into the funnel each time.
6. After all of the solution and rinses have drained through the funnel, use distilled water to slowly rinse the precipitate on the filter paper in the funnel to remove any soluble impurities.
7. Carefully remove the filter paper from the funnel, and place it on the paper towel that you labeled with your name. Unfold the filter paper, and place the paper towel, filter paper, and precipitate in the rinsed beaker. Then, place the beaker in the drying oven. For best results, allow the precipitate to dry overnight.
8. Using beaker tongs, remove your sample from the oven, and let it cool. Record the total mass of the beaker, paper towel, filter paper, and precipitate to the nearest 0.01 g.

FIGURE A The precipitate is a product of the reaction between Na_2CO_3 and $SrCl_2$. Add enough $SrCl_2$ to react with all of the Na_2CO_3 present.

CLEANUP AND DISPOSAL

9. Dispose of the precipitate and the filtrate in designated waste containers. Clean up all equipment after use, and dispose of substances according to your teacher's instructions. Wash your hands thoroughly after all lab work is finished.

ANALYSIS AND INTERPRETATION

1. **Organizing Ideas:** Write a balanced equation for the reaction. What is the precipitate?
2. **Applying Ideas:** Calculate the mass of the dry precipitate. Calculate the number of moles of precipitate produced in the reaction.
3. **Applying Ideas:** How many moles of Na_2CO_3 were present in the 15 mL sample? How many grams of Na_2CO_3 were present?

CONCLUSIONS

1. **Applying Conclusions:** There are 0.30 mol $SrCl_2$ in every liter of solution. Calculate the number of moles of $SrCl_2$ that were added. What is the limiting reactant?

CHAPTER 10

States of Matter

The total three-dimensional arrangement of particles of a crystal is its crystal structure.

Icicle Sculpture

The Kinetic-Molecular Theory of Matter

SECTION 1

OBJECTIVES

- State the kinetic-molecular theory of matter, and describe how it explains certain properties of matter.
- List the five assumptions of the kinetic-molecular theory of gases. Define the terms ideal gas and real gas.
- Describe each of the following characteristic properties of gases: expansion, density, fluidity, compressibility, diffusion, and effusion.
- Describe the conditions under which a real gas deviates from "ideal" behavior.

In Chapter 1, you read that matter exists on Earth in the forms of solids, liquids, and gases. Although it is not usually possible to observe individual particles directly, scientists have studied large groups of these particles as they occur in solids, liquids, and gases.

In the late nineteenth century, scientists developed the kinetic-molecular theory of matter to account for the behavior of the atoms and molecules that make up matter. *The* **kinetic-molecular theory** *is based on the idea that particles of matter are always in motion*. The theory can be used to explain the properties of solids, liquids, and gases in terms of the energy of particles and the forces that act between them. In this section, you will study the theory as it applies to gas molecules.

The Kinetic-Molecular Theory of Gases

The kinetic-molecular theory can help you understand the behavior of gas molecules and the physical properties of gases. The theory provides a model of what is called an ideal gas. *An* **ideal gas** *is a hypothetical gas that perfectly fits all the assumptions of the kinetic-molecular theory.*

The kinetic-molecular theory of gases is based on the following five assumptions:

1. *Gases consist of large numbers of tiny particles that are far apart relative to their size.* These particles, usually molecules or atoms, typically occupy a volume that is about 1000 times greater than the volume occupied by an equal number of particles in the liquid or solid state. Thus, molecules of gases are much farther apart than molecules of liquids or solids. Most of the volume occupied by a gas is empty space, which is the reason that gases have a lower density than liquids and solids do. This also explains the fact that gases are easily compressed.
2. *Collisions between gas particles and between particles and container walls are elastic collisions. An* **elastic collision** *is one in which there is no net loss of total kinetic energy.* Kinetic energy is transferred between two particles during collisions. However, the total kinetic energy of the two particles remains the same as long as temperature is constant.

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Topic: Gases
Code: HC60638

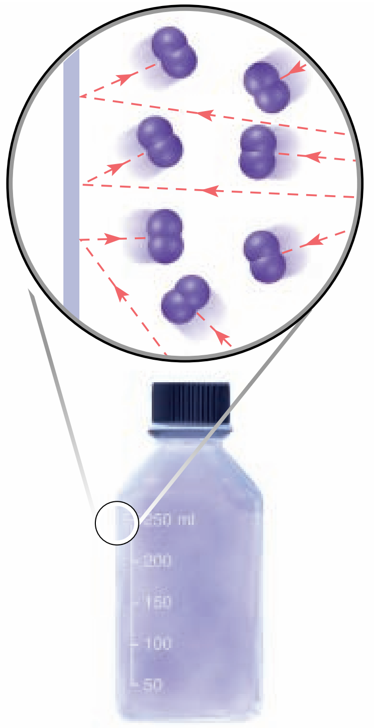

FIGURE 1 Gas particles travel in a straight-line motion until they collide with each other or the walls of their container.

3. *Gas particles are in continuous, rapid, random motion. They therefore possess kinetic energy, which is energy of motion.* Gas particles move in all directions, as shown in **Figure 1.** The kinetic energy of the particles overcomes the attractive forces between them, except near the temperature at which the gas condenses and becomes a liquid.
4. *There are no forces of attraction between gas particles.* You can think of ideal gas molecules as behaving like small billiard balls. When they collide, they do not stick together but immediately bounce apart.
5. *The temperature of a gas depends on the average kinetic energy of the particles of the gas.* The kinetic energy of any moving object, including a particle, is given by the following equation:

$$KE = \frac{1}{2}mv^2$$

In the equation, m is the mass of the particle and v is its speed. Because all the particles of a specific gas have the same mass, their kinetic energies depend only on their speeds. The average speeds and kinetic energies of gas particles increase with an increase in temperature and decrease with a decrease in temperature.

All gases at the same temperature have the same average kinetic energy. Therefore, at the same temperature, lighter gas particles, such as hydrogen molecules, have higher average speeds than do heavier gas particles, such as oxygen molecules.

The Kinetic-Molecular Theory and the Nature of Gases

The kinetic-molecular theory applies only to ideal gases. Although ideal gases do not actually exist, many gases behave nearly ideally if pressure is not very high and temperature is not very low. In the following sections, you will see how the kinetic-molecular theory accounts for the physical properties of gases.

Expansion

Gases do not have a definite shape or a definite volume. They completely fill any container in which they are enclosed, and they take its shape. A gas transferred from a one-liter vessel to a two-liter vessel will quickly expand to fill the entire two-liter volume. The kinetic-molecular theory explains these facts. According to the theory, gas particles move rapidly in all directions (assumption 3) without significant attraction between them (assumption 4).

Fluidity

Because the attractive forces between gas particles are insignificant (assumption 4), gas particles glide easily past one another. This ability to

extension

Chemical Content

Go to **go.hrw.com** for another version of this content. See the chapters "Physical Characteristics of Gases," "Molecular Composition of Gases," and "Liquids and Solids."

Keyword: HC6STMX

flow causes gases to behave as liquids do. Because liquids and gases flow, they are both referred to as *fluids*.

Low Density

The density of a gaseous substance at atmospheric pressure is about 1/1000 the density of the same substance in the liquid or solid state. The reason is that the particles are so much farther apart in the gaseous state (assumption 1).

Compressibility

During compression, the gas particles, which are initially very far apart (assumption 1), are crowded closer together. The volume of a given sample of a gas can be greatly decreased. Steel cylinders containing gases under pressure are widely used in industry. When they are full, such cylinders may contain more than 100 times as many particles of gas as nonpressurized containers of the same size could contain.

Diffusion and Effusion

Gases spread out and mix with one another, even without being stirred. If the stopper is removed from a container of ammonia in a room, ammonia gas will mix uniformly with the air and spread throughout the room. The random and continuous motion of the ammonia molecules (assumption 3) carries them throughout the available space. *Such spontaneous mixing of the particles of two substances caused by their random motion is called* **diffusion.**

Gases diffuse readily into one another and mix together due to the rapid motion of the molecules and the empty space between the molecules. The gas molecules in each of the two flasks in **Figure 2a** continuously move about in separate flasks because the stopcock is closed. When the stopcock is open, the gas molecules continuously diffuse back and forth from one flask to the other through the opening in the stopcock, as shown in **Figure 2b.**

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Topic: Diffusion/Effusion
Code: HC60407

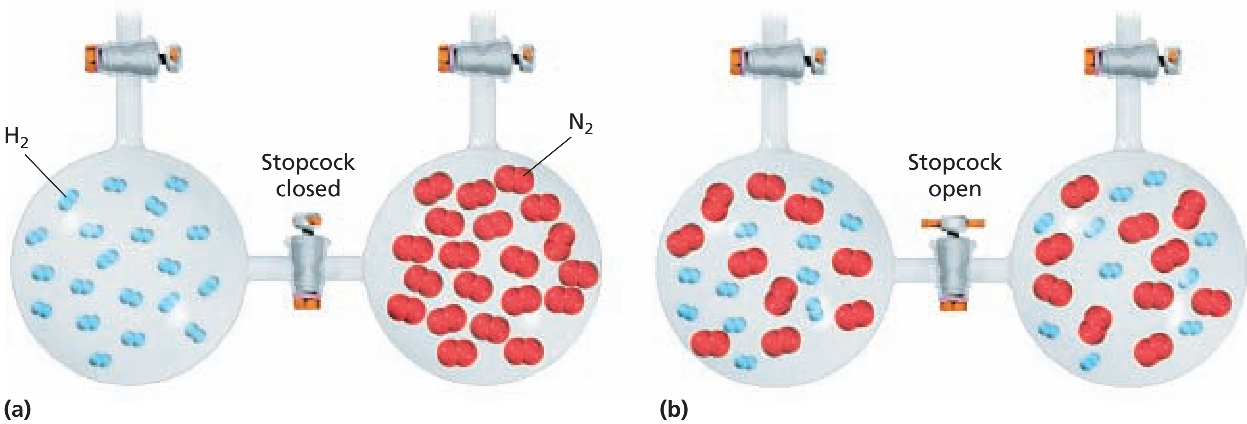

FIGURE 2 Gases diffuse readily into one another. The space between the molecules allows different gases to mix together easily.

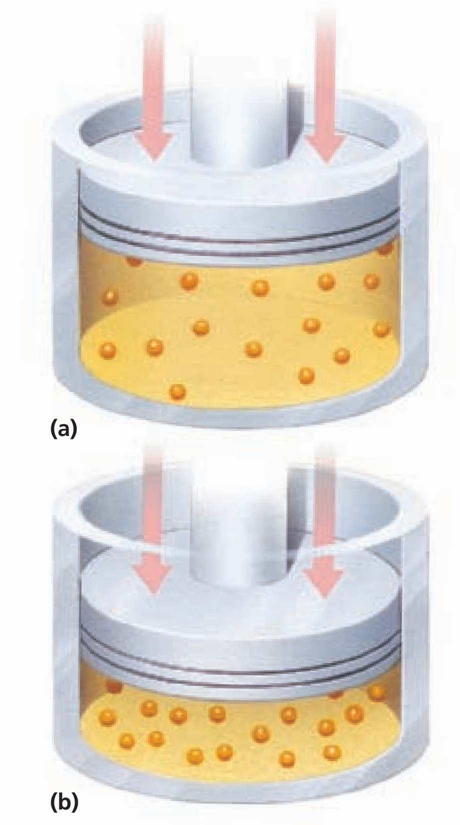

FIGURE 3 (a) Gas molecules in a car engine cylinder expand to fill the cylinder. (b) As pressure is exerted on them, the gas molecules move closer together, reducing their volume.

Diffusion is a process by which particles of a gas spread out spontaneously and mix with other gases. In contrast, **effusion** *is a process by which gas particles pass through a tiny opening*. The rates of effusion of different gases are directly proportional to the velocities of their particles. Because of this proportionality, molecules of low mass effuse faster than molecules of high mass.

Deviations of Real Gases from Ideal Behavior

Because particles of gases occupy space and exert attractive forces on each other, all real gases deviate to some degree from ideal gas behavior. *A* **real gas** *is a gas that does not behave completely according to the assumptions of the kinetic-molecular theory*. At very high pressures and low temperatures, the gas particles will be closer together and their kinetic energy will be insufficient to overcome completely the attractive forces. At such conditions, the gas is most likely to behave like a nonideal gas. These conditions are illustrated in **Figure 3.**

The kinetic-molecular theory is more likely to hold true for gases whose particles have little attraction for each other. The noble gases, such as helium, He, and neon, Ne, show essentially ideal gas behavior over a wide range of temperatures and pressures. The particles of these gases are monatomic and thus nonpolar. The particles of gases, such as nitrogen, N_2, and hydrogen, H_2, are nonpolar diatomic molecules. The behavior of these gases most closely approximates that of the ideal gas under certain conditions. The more polar the molecules of a gas are, the greater the attractive forces between them and the more the gas will deviate from ideal gas behavior. For example, highly polar gases, such as ammonia, NH_3, and water vapor, deviate from ideal behavior to a larger degree than nonpolar gases.

SECTION REVIEW

1. Use the kinetic-molecular theory to explain each of the following properties of gases: expansion, fluidity, low density, compressibility, and diffusion.
2. Describe the conditions under which a real gas is most likely to behave ideally.
3. Which of the following gases would you expect to deviate significantly from ideal behavior: He, O_2, H_2, H_2O, N_2, HCl, or NH_3?
4. How does the kinetic-molecular theory explain the pressure exerted by gases?
5. What happens to gas particles when a gas is compressed?
6. What happens to gas particles when a gas is heated?

Critical Thinking

7. **DRAWING CONCLUSIONS** Molecules of hydrogen escape from Earth, but molecules of oxygen and nitrogen are held to the surface and remain in the atmosphere. Explain.

Liquids

SECTION 2

Objectives

- Describe the motion of particles in liquids and the properties of liquids according to the kinetic-molecular theory.
- Discuss the process by which liquids can change into a gas. Define *vaporization.*
- Discuss the process by which liquids can change into a solid. Define *freezing.*

The water in the waves crashing on a beach and the molten lava rushing down the sides of a volcano are examples of matter in the liquid state. When you think of Earth's oceans, lakes, and rivers and the many liquids you use every day, it is hard to believe that liquids are the *least* common state of matter in the universe. Liquids are less common than solids and gases because a substance can exist in the liquid state only within a relatively narrow range of temperatures and pressures.

In this section, you will examine the properties of the liquid state. You will also compare them with those of the solid state and the gas state. These properties will be discussed in terms of the kinetic-molecular theory.

Properties of Liquids and the Kinetic-Molecular Theory

A liquid can be described as a form of matter that has a definite volume and takes the shape of its container. The properties of liquids can be understood by applying the kinetic-molecular theory, considering the motion and arrangement of molecules and the attractive forces between them.

As in a gas, particles in a liquid are in constant motion. However, the particles in a liquid are closer together than the particles in a gas are. Therefore, the attractive forces between particles in a liquid are more effective than those between particles in a gas. This attraction between liquid particles is caused by the intermolecular forces discussed in Chapter 6: dipole-dipole forces, London dispersion forces, and hydrogen bonding. Some molecules at the surface of a liquid can have enough kinetic energy to overcome these forces, and enter the gas state.

Liquids are more ordered than gases because of the stronger intermolecular forces and the lower mobility of the liquid particles. According to the kinetic-molecular theory of liquids, the particles are not bound together in fixed positions. Instead, they move about constantly. This particle mobility explains why liquids and gases are referred to as fluids. *A* **fluid** *is a substance that can flow and therefore take the shape of its container.* Most liquids naturally flow downhill because of gravity. However, some liquids can flow in other directions as well. For example, liquid helium near absolute zero has the unusual property of being able to flow uphill.

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Topic: Properties of Liquids
Code: HC61228

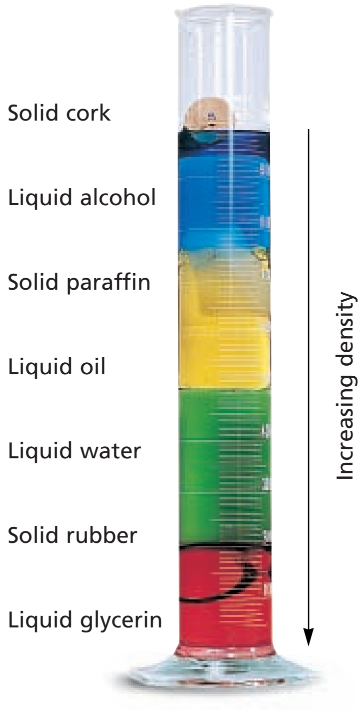


FIGURE 4 Solids and liquids of different densities are shown. The densest materials are at the bottom. The least dense are at the top. (Dyes have been added to the liquids to make the layers more visible.)

Relatively High Density

At normal atmospheric pressure, most substances are hundreds of times denser in a liquid state than in a gaseous state. This higher density is a result of the close arrangement of liquid particles. Most substances are only slightly less dense (about 10%) in a liquid state than in a solid state, however. Water is one of the few substances that becomes less dense when it solidifies, as will be discussed further in Section 5.

At the same temperature and pressure, different liquids can differ greatly in density. **Figure 4** shows some liquids and solids with different densities. The densities differ to such an extent that the liquids form layers.

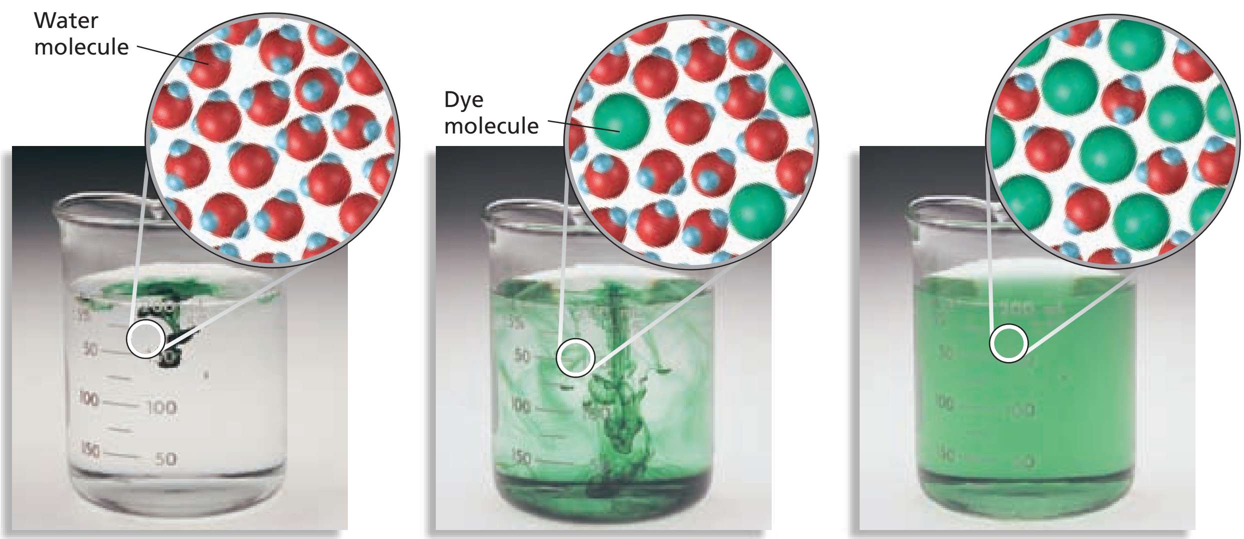
Relative Incompressibility

When liquid water at 20°C is compressed by a pressure of 1000 atm, its volume decreases by only 4%. Such behavior is typical of all liquids and is similar to the behavior of solids. In contrast, a gas under a pressure of 1000 atm would have only about 1/1000 of its volume at normal atmospheric pressure. Liquids are much less compressible than gases because liquid particles are more closely packed together. Like gases, liquids can transmit pressure equally in all directions.

Ability to Diffuse

As described in Section 1, gases diffuse and mix with other gas particles. Liquids also diffuse and mix with other liquids, as shown in **Figure 5.** Any liquid gradually diffuses throughout any other liquid in which it can dissolve. The constant, random motion of particles causes diffusion in liquids, as it does in gases. Yet diffusion is much slower in liquids than in gases because liquid particles are closer together. Also, the attractive forces between the particles of a liquid slow their movement. As the temperature of a liquid is increased, diffusion occurs more rapidly. The reason is that the average kinetic energy, and therefore the average speed of the particles, is increased.

FIGURE 5 Like gases, the two liquids in this beaker diffuse over time. The green liquid food coloring from the drop will eventually form a uniform solution with the water.



Surface Tension

A property common to all liquids is **surface tension,** *a force that tends to pull adjacent parts of a liquid's surface together, thereby decreasing surface area to the smallest possible size.* Surface tension results from the attractive forces between particles of a liquid. The higher the force of attraction, the higher the surface tension. Water has a higher surface tension than most liquids. This is due in large part to the hydrogen bonds water molecules can form with each other. The molecules at the surface of the water are a special case. They can form hydrogen bonds with the other water molecules beneath them and beside them, but not with the molecules in the air above them. As a result, the surface water molecules are drawn together and toward the body of the liquid, creating a high surface tension. Surface tension causes liquid droplets to take on a spherical shape because a sphere has the smallest possible surface area for a given volume. An example of this phenomenon is shown in **Figure 6.**

Capillary action, *the attraction of the surface of a liquid to the surface of a solid,* is a property closely related to surface tension. A liquid will rise quite high in a very narrow tube and will wet the tube if a strong attraction exists between the liquid molecules and the molecules that make up the surface of the tube. This attraction tends to pull the liquid molecules upward along the surface and against the pull of gravity. This process continues until the attractive forces between the liquid molecules and the surface of the tube are balanced by the weight of the liquid. Capillary action can occur between water molecules and paper fibers, as shown in **Figure 7.** Capillary action is at least partly responsible for the transportation of water from the roots of a plant to its leaves. The same process is responsible for the concave liquid surface, called a *meniscus,* that forms in a test tube or graduated cylinder.

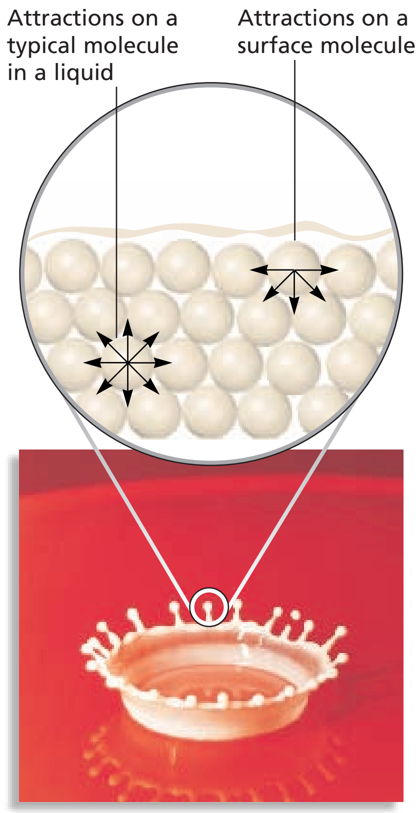
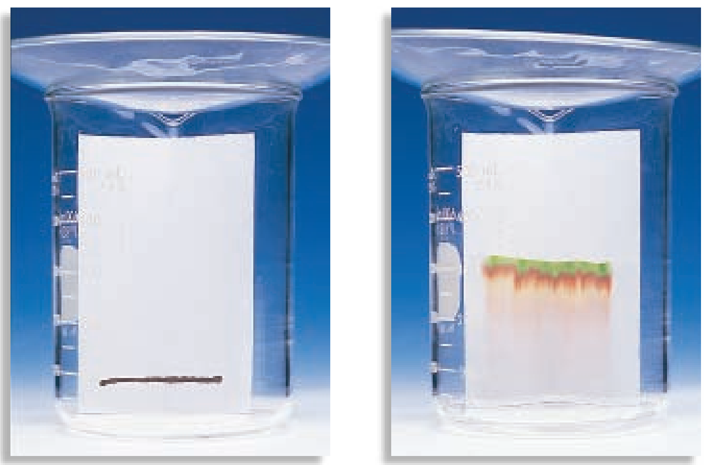


FIGURE 6 As a result of surface tension, liquids form roughly spherical drops. The net attractive forces between the particles pull the molecules on the surface of the drop inward. The molecules are pulled very close together, which minimizes the surface area.

Evaporation and Boiling

The process by which a liquid or solid changes to a gas is **vaporization.** Evaporation is a form of vaporization. **Evaporation** *is the process by which particles escape from the surface of a nonboiling liquid and enter the gas state.*



(a) (b)

FIGURE 7 The attraction between polar water molecules and polar cellulose molecules in paper fibers causes the water to move up in the paper. The water-soluble ink placed near the bottom of the paper in (a) rises up the paper along with the water, as seen in (b). As the ink moves up the paper, it is separated into its various components, producing the different bands of color. This separation occurs because the water and the paper attract the molecules of the ink components differently. These phenomena are used in the separation process of paper chromatography seen here.

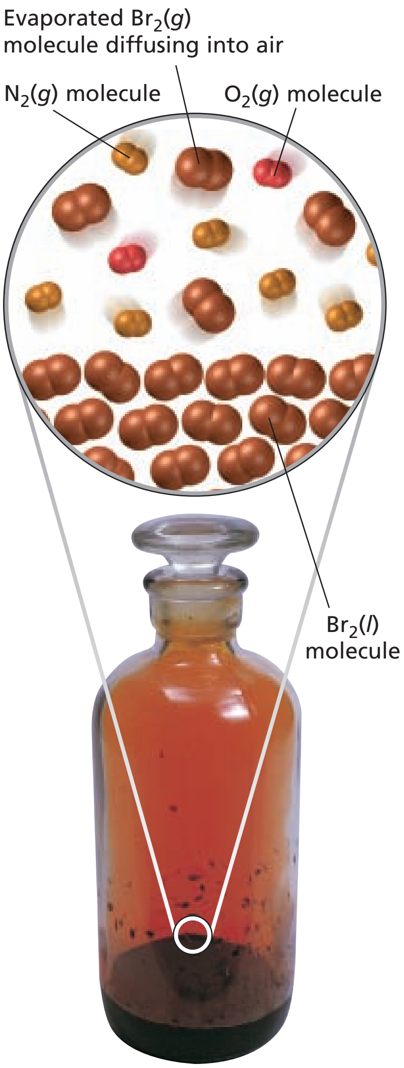


FIGURE 8 Liquid bromine, Br_2, evaporates near room temperature. The resulting brownish red gas diffuses into the air above the surface of the liquid.

A small amount of liquid bromine was added to the bottle shown in **Figure 8.** Within a few minutes, the air above the liquid bromine turned brownish-red because some bromine molecules escaped from the surface of the liquid. These molecules became gas molecules, or bromine vapor, which mixed with the air. A similar phenomenon occurs if you apply perfume to your wrist. Within seconds, you become aware of the perfume's fragrance. Scent molecules evaporate from your skin and diffuse through the air, where your nose detects them.

Evaporation occurs because the particles of a liquid have different kinetic energies. Particles with higher-than-average energies move faster. Some surface particles with higher-than-average energies can overcome the intermolecular forces that bind them to the liquid. They can then escape into the gas state.

Evaporation is a crucial process in nature. Evaporation removes fresh water from the surface of the ocean, leaving behind a higher concentration of salts. In tropical areas, evaporation occurs at a higher rate, causing the surface water to be saltier. All water that falls to Earth in the form of rain and snow previously evaporated from oceans, lakes, and rivers. Evaporation of perspiration plays an important role in keeping you cool. Perspiration, which is mostly water, cools you by absorbing body heat when it evaporates. Energy as heat is absorbed from the skin, causing the cooling effect.

Boiling is the change of a liquid to bubbles of vapor that appear throughout the liquid. Boiling differs from evaporation, as you will see in Section 4.

Formation of Solids

When a liquid is cooled, the average energy of its particles decreases. If the energy is low enough, attractive forces pull the particles into an even more orderly arrangement. The substance then becomes a solid. *The physical change of a liquid to a solid by removal of energy as heat is called* **freezing** *or solidification.* Perhaps the best-known example of freezing is the change of liquid water to solid water, or ice, at 0°C. Another familiar example is the solidification of paraffin at room temperature. All liquids freeze, although not necessarily at temperatures you normally encounter. Ethanol, for example, freezes near –114°C.

SECTION REVIEW

1. Describe the liquid state according to the kinetic-molecular theory.
2. List the properties of liquids.
3. How does the kinetic-molecular theory explain the following properties of liquids: (a) relatively high density, (b) ability to diffuse, and (c) ability to evaporate?
4. Explain why liquids in a test tube form a meniscus.
5. Compare vaporization and evaporation.

Critical Thinking

6. **INTERPRETING CONCEPTS** The evaporation of liquid water from the surface of Earth is an important step in the water cycle. How do water molecules obtain enough kinetic energy to escape into the gas state?

Solids

SECTION 3

OBJECTIVES

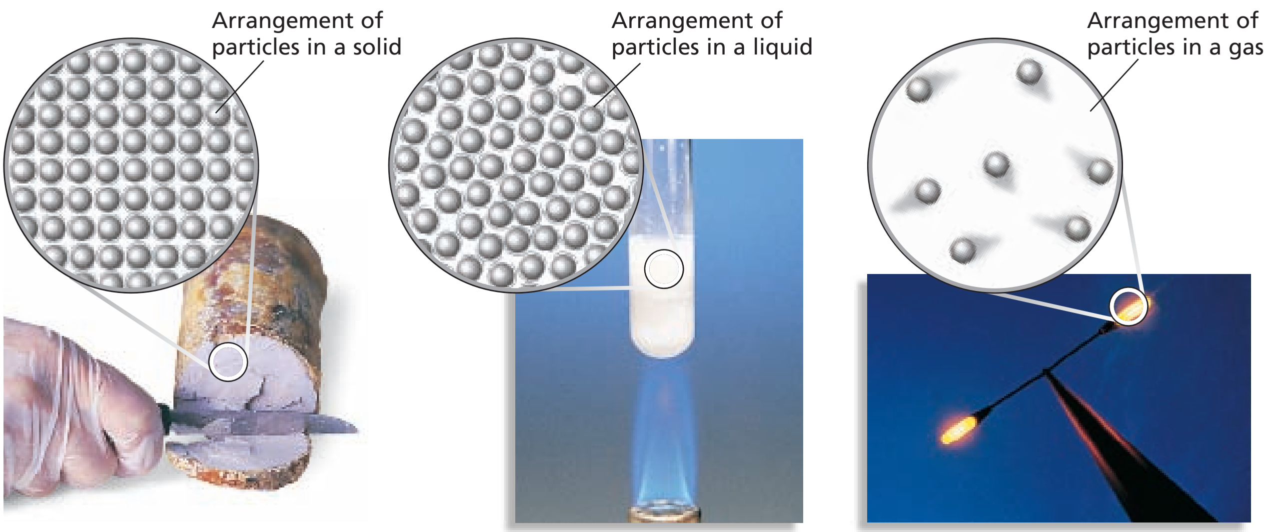
- Describe the motion of particles in solids and the properties of solids according to the kinetic-molecular theory.
- Distinguish between the two types of solids.
- Describe the different types of crystal symmetry. Define *crystal structure* and *unit cell.*

The common expression "solid as a rock" suggests that something is hard or unyielding and has a definite shape and volume. In this section you will examine the properties of solids and compare them with those of liquids and gases. The properties of solids are explained in terms of the kinetic-molecular theory, as the other states of matter are.

Properties of Solids and the Kinetic-Molecular Theory

The particles of a solid are more closely packed than those of a liquid or gas. Intermolecular forces between particles are therefore much more effective in solids. All interparticle attractions such as dipole-dipole attractions, London dispersion forces, and hydrogen bonding exert stronger effects in solids than in the corresponding liquids or gases. Attractive forces tend to hold the particles of a solid in relatively fixed positions, with only vibrational movement around fixed points. Because the motions of the particles are restricted in this way, solids are more ordered than liquids and are much more ordered than gases. The importance of order and disorder in physical and chemical changes will be discussed in Chapter 16. Compare the physical appearance and molecular arrangement of the element in **Figure 9** in solid, liquid, and gas form.

FIGURE 9 Particles of sodium metal in three different states are shown. Sodium exists in a gaseous state in a sodium-vapor lamp.



There are two types of solids: crystalline solids and amorphous solids. Most solids are **crystalline solids**—*they consist of crystals. A* **crystal** *is a substance in which the particles are arranged in an orderly, geometric, repeating pattern.* Noncrystalline solids, including glass and plastics, are called amorphous solids. *An* **amorphous solid** *is one in which the particles are arranged randomly.* The two types of solids will be discussed in more detail later in this section.

Definite Shape and Volume

Unlike liquids and gases, solids can maintain a definite shape without a container. In addition, crystalline solids are geometrically regular. Even the fragments of a shattered crystalline solid have distinct geometric shapes that reflect their internal structure. Amorphous solids maintain a definite shape, but they do not have the distinct geometric shapes of crystalline solids. For example, glass can be molded into any shape. If it is shattered, glass fragments can have a wide variety of irregular shapes.

The volume of a solid changes only slightly with a change in temperature or pressure. Solids have definite volume because their particles are packed closely together. There is very little empty space into which the particles can be compressed. Crystalline solids generally do not flow because their particles are held in relatively fixed positions.

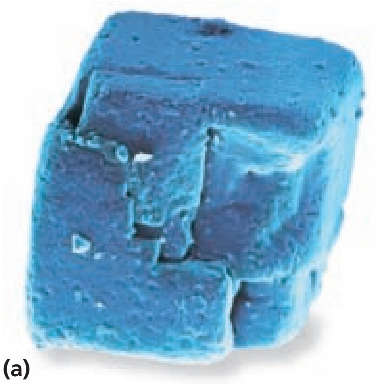
Definite Melting Point

Melting *is the physical change of a solid to a liquid by the addition of energy as heat. The temperature at which a solid becomes a liquid is its* **melting point.** At this temperature, the kinetic energies of the particles within the solid overcome the attractive forces holding them together. The particles can then break out of their positions in crystalline solids, which have definite melting points. In contrast, amorphous solids, such as glass and plastics, have no definite melting point. They have the ability to flow over a range of temperatures. Therefore, amorphous solids are sometimes classified as **supercooled liquids,** *which are substances that retain certain liquid properties even at temperatures at which they appear to be solid.* These properties exist because the particles in amorphous solids are arranged randomly, much like the particles in a liquid. Unlike the particles in a true liquid, however, the particles in amorphous solids are not constantly changing their positions.

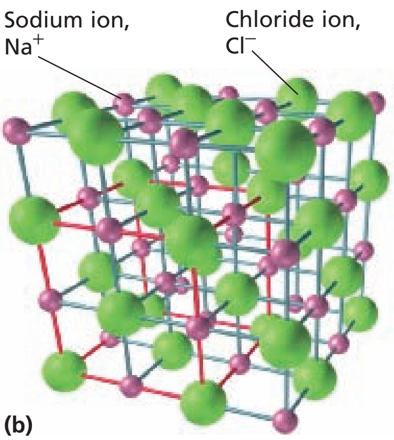
High Density and Incompressibility

In general, substances are most dense in the solid state. Solids tend to be slightly denser than liquids and much denser than gases. The higher density results from the fact that the particles of a solid are more closely packed than those of a liquid or a gas. Solid hydrogen is the least dense solid; it has a density of about 1/320 of the densest element, osmium, Os.

Solids are generally less compressible than liquids. For practical purposes, solids can be considered incompressible. Some solids, such as wood and cork, may *seem* compressible, but they are not. They contain pores that are filled with air. When subjected to intense pressure, the pores are compressed, not the solid matter in the wood or cork itself.

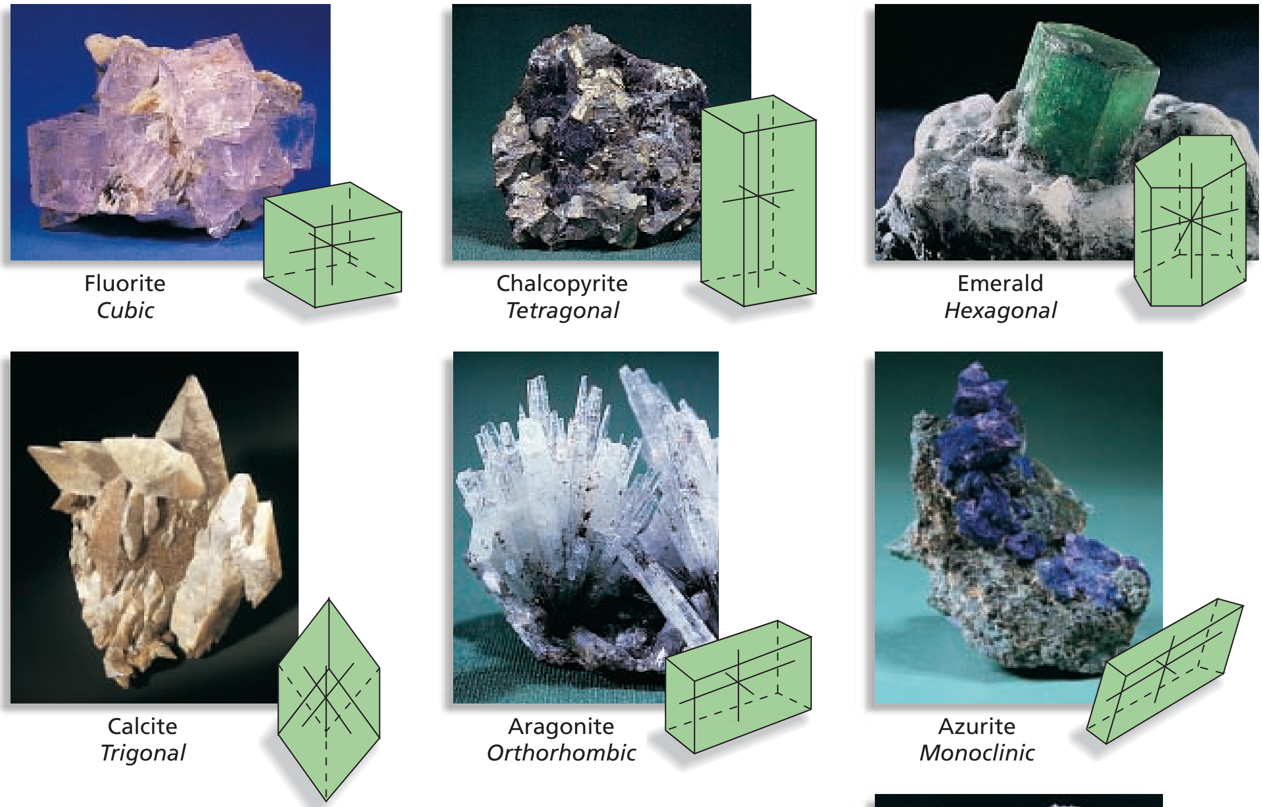


(a)



(b)

FIGURE 10 (a) This is a scanning electron micrograph (SEM) of a sodium chloride crystal. A sodium chloride crystal can be represented by its crystal structure (b), which is made up of individual unit cells represented regularly in three dimensions. Here, one unit cell is outlined in red.



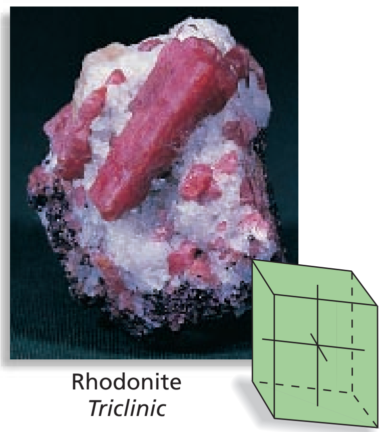


FIGURE 11 Shown are the seven basic crystalline systems and representative minerals of each.

Low Rate of Diffusion

If a zinc plate and a copper plate are clamped together for a long time, a few atoms of each metal will diffuse into the other. This observation shows that diffusion does occur in solids. The rate of diffusion is millions of times slower in solids than in liquids, however.

Crystalline Solids

Crystalline solids exist either as single crystals or as groups of crystals fused together. *The total three-dimensional arrangement of particles of a crystal is called a* **crystal structure.** The arrangement of particles in the crystal can be represented by a coordinate system called a *lattice. The smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice is called a* **unit cell.** Each crystal lattice contains many unit cells packed together. **Figure 10** shows the relationship between a crystal lattice and its unit cell. A crystal and its unit cells can have any one of seven types of symmetry. This fact enables scientists to classify crystals by their shape. Diagrams and examples of each type of crystal symmetry are shown in **Figure 11.**

Binding Forces in Crystals

Crystal structures can also be described in terms of the types of particles in them and the types of chemical bonding between the particles.

TABLE 1 ***Melting and Boiling Points of Representative Crystalline Solids***

Type of substance	Formula	Melting point (°C)	Boiling point at 1 atm (°C)
Ionic	NaCl	801	1413
	MgF_2	1266	2239
Covalent network	$(SiO_2)_x$	1610	2230
	C_x (diamond)	3500	3930
Metallic	Hg	–39	357
	Cu	1083	2567
	Fe	1535	2750
	W	3410	5660
Covalent molecular (nonpolar)	H_2	–259	–253
	O_2	–218	–183
	CH_4	–182	–164
	CCl_4	–23	77
	C_6H_6	6	80
Covalent molecular (polar)	NH_3	–78	–33
	H_2O	0	100

According to this method of classification, there are four types of crystals. These types are listed in **Table 1.** Refer to this table as you read the following discussion.

1. *Ionic crystals.* The ionic crystal structure consists of positive and negative ions arranged in a regular pattern. The ions can be monatomic or polyatomic. Generally, ionic crystals form when Group 1 or Group 2 metals combine with Group 16 or Group 17 nonmetals or nonmetallic polyatomic ions. The strong binding forces between the positive and negative ions in the crystal structure give the ionic crystals certain properties. For example, these crystals are hard and brittle, have high melting points, and are good insulators.
2. *Covalent network crystals.* In covalent network crystals, each atom is covalently bonded to its nearest neighboring atoms. The covalent bonding extends throughout a network that includes a very large number of atoms. Three-dimensional covalent network solids include diamond, C_x, quartz, $(SiO_2)_x$—shown in **Figure 12**—silicon carbide, $(SiC)_x$, and many oxides of transition metals. Such solids are essentially giant molecules. The subscript x in these formulas indicates that the component within the parentheses extends indefinitely. The network solids are nearly always very hard and brittle. They have rather high melting points and are usually nonconductors or semiconductors.
3. *Metallic crystals.* The metallic crystal structure consists of metal cations surrounded by a sea of delocalized valence electrons. The electrons come from the metal atoms and belong to the crystal as a whole. The freedom of these delocalized electrons to move throughout the crystal explains the high electric conductivity of metals.

4. *Covalent molecular crystals.* The crystal structure of a covalent molecular substance consists of covalently bonded molecules held together by intermolecular forces. If the molecules are nonpolar—for example, hydrogen, H_2, methane, CH_4, and benzene, C_6H_6—then there are only weak London dispersion forces between molecules. In a polar covalent molecular crystal—for example, water, H_2O, and ammonia, NH_3—molecules are held together by dispersion forces, by somewhat stronger dipole-dipole forces, and sometimes by even stronger hydrogen bonding. The forces that hold polar or nonpolar molecules together in the structure are much weaker than the covalent chemical bonds between the atoms within each molecule. Covalent molecular crystals thus have low melting points. They are easily vaporized, are relatively soft, and are good insulators. Ice crystals, the most familiar molecular crystals, are discussed in Section 5.

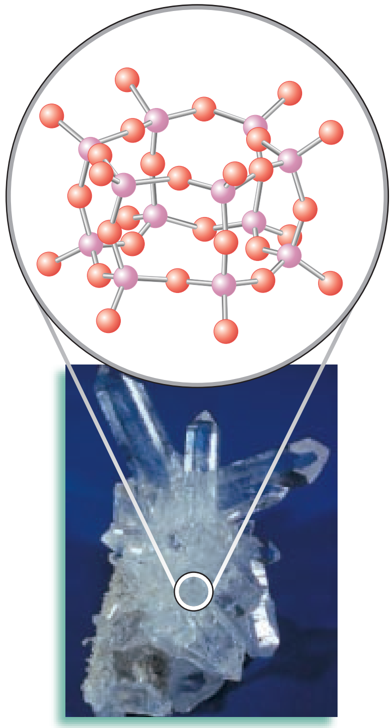


FIGURE 12 Covalent network crystals include three-dimensional network solids, such as this quartz, $(SiO_2)_x$, shown here with its three-dimensional atomic structure.

Amorphous Solids

The word *amorphous* comes from the Greek for "without shape." Unlike the atoms that form crystals, the atoms that make up amorphous solids, such as glasses and plastics, are not arranged in a regular pattern.

Glasses are made by cooling certain molten materials in a way that prevents them from crystallizing. The properties that result make glasses suitable for many uses, including windows, light bulbs, transformer cores, and optical fibers that carry telephone conversations.

Plastics, another type of amorphous solid, are easily molded at high temperatures and pressures. They are used in many structural materials.

Other, more recently created amorphous solids have been placed in many important applications. Amorphous semiconductors are used in electronic devices, including solar cells, copiers, laser printers, and flat-panel displays for computer monitors and television screens.

SECTION REVIEW

1. Describe the solid state according to the kinetic-molecular theory.
2. What is the difference between an amorphous solid and a crystalline solid?
3. Account for each of the following properties of solids: (a) the definite volume, (b) the relatively high density of solids, (c) the extremely low rate of diffusion.
4. Compare and contrast the four types of crystals.
5. Why do crystalline solids shatter into regularly shaped fragments when broken?

Critical Thinking

6. **RELATING IDEAS** Explain why ionic crystals melt at much higher temperatures than typical covalent molecular crystals?

SECTION 4

Changes of State

OBJECTIVES

- Explain the relationship between equilibrium and changes of state.
- Interpret phase diagrams.
- Explain what is meant by equilibrium vapor pressure.
- Describe the processes of boiling, freezing, melting, and sublimation.

Matter on Earth can exist in any of these states—gas, liquid, or solid—and can change from one state to another. **Table 2** lists the possible changes of state. In this section, you will examine these changes of state and the factors that determine them.

Changes of State and Equilibrium

Some liquid chemical substances, such as rubbing alcohol, have an odor that is very easily detected. This is because some molecules at the upper surface of the liquid have enough energy to overcome the attraction of neighboring molecules. These molecules leave the liquid phase and evaporate. *A* **phase** *is any part of a system that has uniform composition and properties.* In a closed bottle of rubbing alcohol, the gas molecules are confined to the area under the cap. Some of the gas molecules strike the liquid surface and reenter the liquid phase through *condensation.* **Condensation** *is the process by which a gas changes to a liquid.* A gas in contact with its liquid or solid phase is often called a *vapor.*

If the temperature of the liquid remains constant and the cap remains closed, the rate at which molecules move from the liquid phase to the vapor phase remains constant. Near the beginning of the evaporation process, very few molecules are in the gas phase, so the rate of condensation is very low. But as more liquid evaporates, the increasing number of gas molecules causes the rate of condensation to increase. Eventually, the rate of condensation equals the rate of evaporation, and a state of equilibrium is established, as shown in **Figure 13. Equilibrium** *is a dynamic condition in which two opposing changes occur at equal rates in a closed system.* Even though molecules are constantly moving

TABLE 2 Possible Changes of State

Change of state	Process	Example
Solid $\longrightarrow$ liquid	melting	ice $\longrightarrow$ water
Solid $\longrightarrow$ gas	sublimation	dry ice $\longrightarrow$ CO_2 gas
Liquid $\longrightarrow$ solid	freezing	water $\longrightarrow$ ice
Liquid $\longrightarrow$ gas	vaporization	liquid bromine $\longrightarrow$ bromine vapor
Gas $\longrightarrow$ liquid	condensation	water vapor $\longrightarrow$ water
Gas $\longrightarrow$ solid	deposition	water vapor $\longrightarrow$ ice

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Code: HC60561

FIGURE 13 A liquid-vapor equilibrium develops in a closed system. (a) At first there is only liquid present, but molecules are beginning to evaporate. (b) Evaporation continues at a constant rate. Some vapor molecules are beginning to condense to liquid. (c) Equilibrium has been reached between the rate of condensation and the rate of evaporation.

between liquid and gas phases, there is no net change in the amount of substance in either phase.

Equilibrium Vapor Pressure of a Liquid

Vapor molecules in equilibrium with a liquid in a closed system exert a pressure proportional to the concentration of molecules in the vapor phase. *The pressure exerted by a vapor in equilibrium with its corresponding liquid at a given temperature is called the* **equilibrium vapor pressure** *of the liquid.*

The increase in equilibrium vapor pressure with increasing temperature can be explained in terms of the kinetic-molecular theory for the liquid and gaseous states. Increasing the temperature of a liquid increases the average kinetic energy of the liquid's molecules. This energy change increases the number of molecules that have enough energy to escape from the liquid phase into the vapor phase. The resulting increased evaporation rate increases the number of molecules in the vapor phase, which in turn increases the equilibrium vapor pressure.

Because all liquids have characteristic forces of attraction between their particles, every liquid has a specific equilibrium vapor pressure at a given temperature. The stronger these attractive forces are, the smaller the percentage of liquid particles that can evaporate at any given temperature is. A low percentage of evaporation results in a low equilibrium vapor pressure is. **Volatile liquids,** *which are liquids that evaporate readily,* have relatively weak forces of attraction between their particles. Ether is a typical volatile liquid. Nonvolatile liquids do not evaporate readily, and have relatively strong attractive forces between their particles. Molten ionic compounds are examples of nonvolatile liquids.

extension

Chemical Content

Go to **go.hrw.com** for more information on equilibrium and changes of state.

Keyword: HC6STMX

Boiling

Equilibrium vapor pressures can be used to explain and define the concept of boiling, which you read about in Section 3. **Boiling** *is the conversion of a liquid to a vapor within the liquid as well as at its surface.*

If the temperature of the liquid is increased, the equilibrium vapor pressure also increases. *The* **boiling point** *of a liquid is the temperature at which the equilibrium vapor pressure of the liquid equals the atmospheric pressure.* The lower the atmospheric pressure is, the lower the boiling point is.

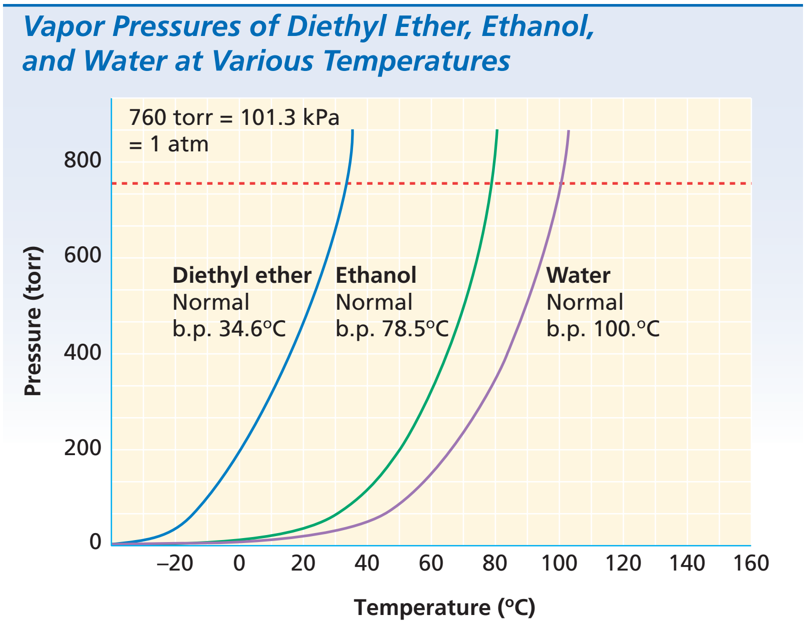


FIGURE 14 The vapor pressure of any liquid increases as its temperature increases. A liquid boils when its vapor pressure equals the pressure of the atmosphere.

At the boiling point, all of the energy absorbed is used to evaporate the liquid, and the temperature remains constant as long as the pressure does not change. If the pressure above the liquid being heated is increased, the temperature of the liquid will rise until the vapor pressure equals the new pressure and the liquid boils once again. This is the principle behind the operation of a pressure cooker. The cooker is sealed so that steam pressure builds up over the surface of the boiling water inside. The pressure increases the boiling temperature of the water, resulting in shorter cooking times. Conversely, a device called a *vacuum evaporator* causes boiling at lower-than-normal temperatures. Vacuum evaporators are used to remove water from milk and sugar solutions. Under reduced pressure, the water boils away at a temperature low enough to avoid scorching the milk or sugar. This process is used to prepare evaporated milk and sweetened condensed milk.

At normal atmospheric pressure (1 atm, 760 torr, or 101.3 kPa), the boiling point of water is exactly 100°C. This temperature is known as the *normal* boiling point of water. **Figure 14** shows that the normal boiling point of each liquid is the temperature at which the liquid's equilibrium vapor pressure equals 760 torr.

Energy and Boiling

Energy must be added continuously in order to keep a liquid boiling. A pot of boiling water stops boiling almost immediately after it is removed from the stove. If you were to carefully measure the temperature of a boiling liquid and its vapor you would find that they are at the same constant temperature. The temperature at the boiling point remains constant despite the continuous addition of energy. The added energy is used to overcome the attractive forces between molecules of the liquid during the liquid-to-gas change and is stored in the vapor as potential energy.

Energy Distribution of Molecules in a Liquid at Different Temperatures

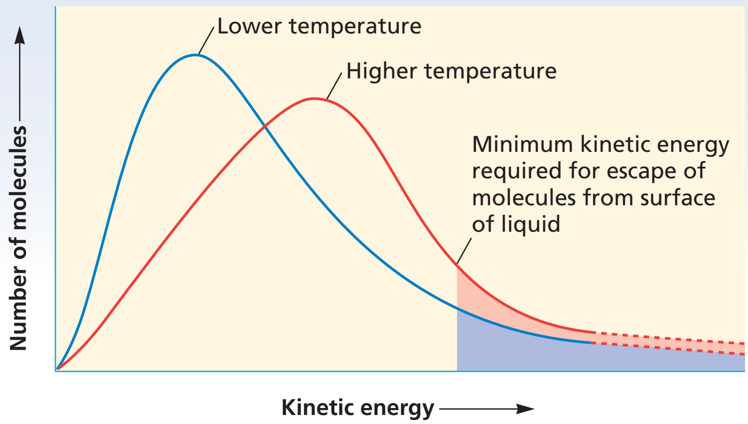


FIGURE 15 The number of molecules in a liquid with various kinetic energies is represented at two different temperatures. Notice the shaded area, which shows the fraction of the molecules that have at least the minimum amount of kinetic energy required for evaporation to take place.

Molar Enthalpy of Vaporization

The amount of energy as heat that is needed to vaporize one mole of liquid at the liquid's boiling point at constant pressure is called the liquid's **molar enthalpy of vaporization,** ΔH_v. The magnitude of the molar enthalpy of vaporization is a measure of the attraction between particles of the liquid. The stronger this attraction is, the more energy that is required to overcome it, which results in a higher molar enthalpy of vaporization. Each liquid has a characteristic molar enthalpy of vaporization. Compared with other liquids, water has an unusually high molar enthalpy of vaporization due to the extensive hydrogen bonding in liquid water. This property makes water a very effective cooling agent. When water evaporates from your skin, the escaping molecules carry a great deal of energy as heat away with them. **Figure 15** shows the distribution of the kinetic energies of molecules in a liquid at two different temperatures. You can see that at the higher temperature, a greater portion of the molecules have the kinetic energy required to escape from the liquid surface and become vapor.

Freezing and Melting

As you learned in Section 2, *the physical change of a liquid to a solid is called* **freezing.** Freezing involves a loss of energy in the form of heat by the liquid and can be represented by the following reaction.

$$\text{liquid} \longrightarrow \text{solid} + \text{energy}$$

In the case of a pure crystalline substance, this change occurs at constant temperature. *The normal* **freezing point** *is the temperature at which the solid and liquid are in equilibrium at 1 atm (760 torr, or 101.3 kPa) pressure.* At the freezing point, particles of the liquid and the solid have the same average kinetic energy. Therefore, the energy loss during freezing is a loss of potential energy that was present in the liquid. At the same time energy decreases, there is a significant increase in particle order because the solid state of a substance is much more ordered than the liquid state, even at the same temperature.

Melting, the reverse of freezing, also occurs at constant temperature. As a solid melts, it continuously absorbs energy as heat, as represented by the following equation.

$$\text{solid} + \text{energy} \longrightarrow \text{liquid}$$

For pure crystalline solids, the melting point and freezing point are the same. At equilibrium, melting and freezing proceed at equal rates. The following general equilibrium equation can be used to represent these states.

$$\text{solid} + \text{energy} \rightleftarrows \text{liquid}$$

At normal atmospheric pressure, the temperature of a system containing ice and liquid water will remain at 0.°C as long as both ice and water are present. That temperature will persist no matter what the surrounding temperature. Adding energy in the form of heat to such a system shifts the equilibrium to the right. That shift increases the proportion of liquid water and decreases that of ice. Only after all the ice has melted will the addition of energy increase the temperature of the system.

Chemistry in Action

Surface Melting

Freezing of water and melting of ice are phase changes that are familiar to all of us. Yet physicists and chemists have only recently begun to understand the basic aspects of these phase changes, with experimental and theoretical studies of a phenomenon known as surface melting. Experimental studies in the mid-1980s confirmed that the rigid surface arrangements of metals can become increasingly disordered several degrees below the melting point of the metal, forming a "quasi-liquid layer." Many different techniques have now shown that ice also has such a fluid surface layer just a few molecules thick. This surface melting of ice might explain observations as diverse as the origin of lightning, the unique shapes of snowflakes, and ice skating.

Molar Enthalpy of Fusion

The amount of energy as heat required to melt one mole of solid at the solid's melting point is the solid's **molar enthalpy of fusion,** ΔH_f. The energy absorbed increases the solid's potential energy as its particles are pulled apart, overcoming the attractive forces holding them together. At the same time, there is a significant decrease in particle order as the substance makes the transformation from solid to liquid. Similar to the molar enthalpy of vaporization, the magnitude of the molar enthalpy of fusion depends on the attraction between the solid particles.

Sublimation and Deposition

At sufficiently low temperature and pressure conditions, a liquid cannot exist. Under such conditions, a solid substance exists in equilibrium with its vapor instead of its liquid, as represented by the following equation.

$$\text{solid} + \text{energy} \rightleftarrows \text{vapor}$$

The change of state from a solid directly to a gas is known as **sublimation.** The reverse process is called **deposition,** *the change of state from a gas directly to a solid.* Among the common substances that sublime at ordinary temperatures are dry ice (solid CO_2) and iodine. Ordinary ice sublimes slowly at temperatures lower than its melting point (0.°C). This explains how a thin layer of snow can eventually disappear, even if the temperature remains below 0.°C. Sublimation occurs in frost-free refrigerators when the temperature in the freezer compartment is periodically raised to cause any ice that has formed to sublime. A blower then removes the water vapor that has formed. The formation of frost on a cold surface is a familiar example of deposition.

Phase Diagrams

A **phase diagram** *is a graph of pressure versus temperature that shows the conditions under which the phases of a substance exist.* A phase diagram also reveals how the states of a system change with changing temperature or pressure.

Figure 16 shows the phase diagram for water over a range of temperatures and pressures. Note the three curves, AB, AC, and AD. Curve AB indicates the temperature and pressure conditions at which ice and water vapor can coexist at equilibrium. Curve AC indicates the temperature and pressure conditions at which liquid water and water vapor coexist at equilibrium. Similarly, curve AD indicates the temperature and pressure conditions at which ice and liquid water coexist at equilibrium. Because ice is less dense than liquid water, an increase in pressure lowers the melting point. (Most substances have a positive slope for this curve.) Point A is the triple point of water. *The* **triple point** *of a substance indicates the temperature and pressure conditions at which the solid, liquid, and vapor of the substance can coexist at equilibrium.* Point C is the critical point of water. *The* **critical point** *of a substance indicates the critical temperature and critical pressure. The* **critical temperature** *(t_c) is the temperature above which the substance cannot exist in the liquid state.* The critical temperature of water is 373.99°C. Above this

extension

CROSS-DISCIPLINARY

Go to **go.hrw.com** for a full-length article on how cloud seeding is used to produce precipitation.

Keyword: HC6STMX

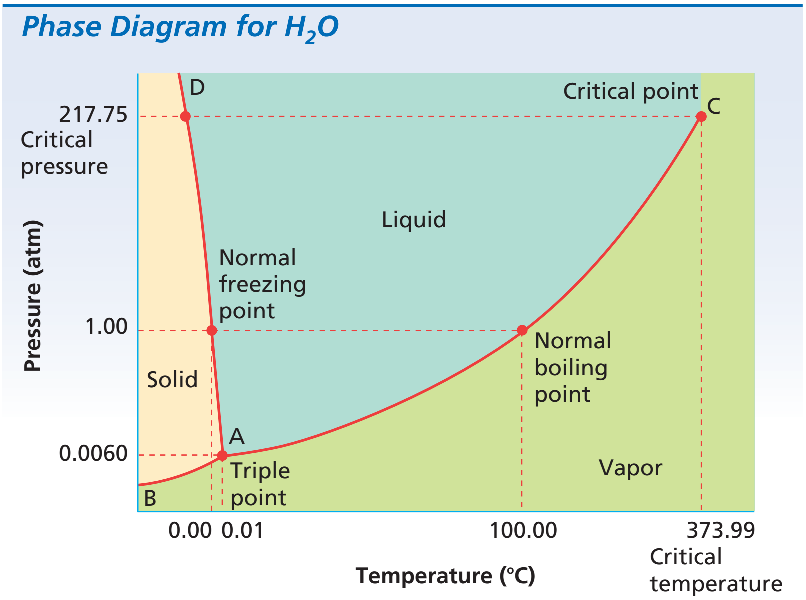
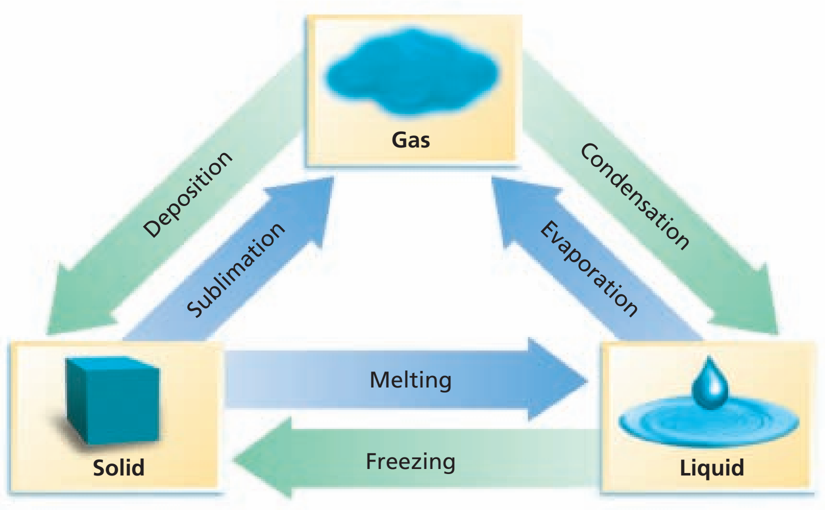


FIGURE 16 This phase diagram shows the relationships between the physical states of water and its pressure and temperature.

FIGURE 17 Solids, liquids, and gases can undergo various changes of state. The changes shown in green are exothermic, and those shown in blue are endothermic.



extension

Chemistry in Action

Go to **go.hrw.com** for a full-length article on phase-change materials.

Keyword: HC6STMX

temperature, water cannot be liquefied, no matter how much pressure is applied. *The* **critical pressure** *(P_c) is the lowest pressure at which the substance can exist as a liquid at the critical temperature.* The critical pressure of water is 217.75 atm.

The phase diagram in **Figure 16** indicates the normal boiling point and the normal freezing point of water. It also shows how boiling point and freezing point change with pressure. As shown by the slope of line AD, ice melts at a higher temperature with decreasing pressure. Below the triple point, the temperature of sublimation decreases with decreasing pressure. Foods are freeze-dried by freezing the food and then lowering the pressure to cause the ice in the food to sublime rather than melt. **Figure 17** summarizes the changes of state of solids, liquids, and gases.

SECTION REVIEW

1. What is equilibrium?
2. What happens when a liquid-vapor system at equilibrium experiences an increase in temperature? What happens when it experiences a decrease in temperature?
3. What would be an example of deposition?
4. What is the equilibrium vapor pressure of a liquid? How is it measured?
5. What is the boiling point of a liquid?
6. In the phase diagram for water, what is meant by the triple point and the critical point?

Critical Thinking

7. **INTERPRETING GRAPHICS** Refer to the phase diagram for water on page 347 to answer the following questions.
   a. Describe all the changes a sample of solid water would undergo when heated from −10°C to its critical temperature at a pressure of 1.00 atm.
   b. Describe all the changes a sample of water vapor would undergo when cooled from 110°C to 5°C at a pressure of 1.00 atm.
   c. At approximately what pressure will water be a vapor at 0°C?
   d. Within what range of pressures will water be a liquid at temperatures above its normal boiling point?

Water

SECTION 5

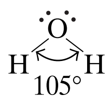
Objectives

- Describe the structure of a water molecule.
- Discuss the physical properties of water. Explain how they are determined by the structure of water.
- Calculate the amount of energy absorbed or released when a quantity of water changes state.

Water is a familiar substance in all three physical states: solid, liquid, and gas. On Earth, water is by far the most abundant liquid. Oceans, rivers, and lakes cover about 75% of Earth's surface. Significant quantities of water are also frozen in glaciers. Water is an essential component of all organisms; 70% to 90% of the mass of living things is water. The chemical reactions of most life processes take place in water, and water is frequently a reactant or product in such reactions. In order to better understand the importance of water, let us take a closer look at its structure and its properties.

Structure of Water

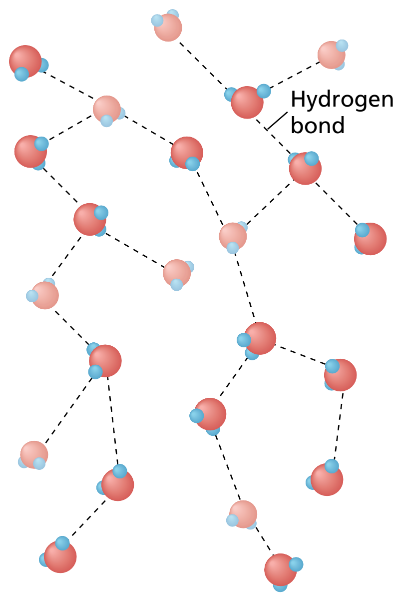
As discussed in Chapter 6, water molecules consist of two atoms of hydrogen and one atom of oxygen united by polar-covalent bonds. Research shows that a water molecule is bent. The structure can be represented as follows.



The angle between the two hydrogen-oxygen bonds is about 105°. This is close to the angle expected for sp^3 hybridization of the oxygen-atom orbitals.

The molecules in solid or liquid water are linked by hydrogen bonding. The number of linked molecules decreases with increasing temperature because increases in kinetic energy make hydrogen bond formation difficult. Nevertheless, there are usually from four to eight molecules per group in liquid water, as shown in **Figure 18.** If it were not for these molecular groups, water would be a gas at room temperature. Nonpolar molecules, such as methane, CH_4, that are similar in size and mass to water molecules do not undergo hydrogen bonding. Such substances are gases at room temperature.

Ice consists of water molecules in the hexagonal arrangement shown in **Figure 19.** The empty spaces between molecules in this pattern account for the relatively low density of ice. As ice is heated, the increased energy of the molecules causes them to move and vibrate more vigorously. When the melting point is reached, the energy of the



Liquid water

FIGURE 18 The structure of liquid water shows that within the water molecule, oxygen and hydrogen are covalently bonded to each other, while the molecules are held together in groups by hydrogen bonds.

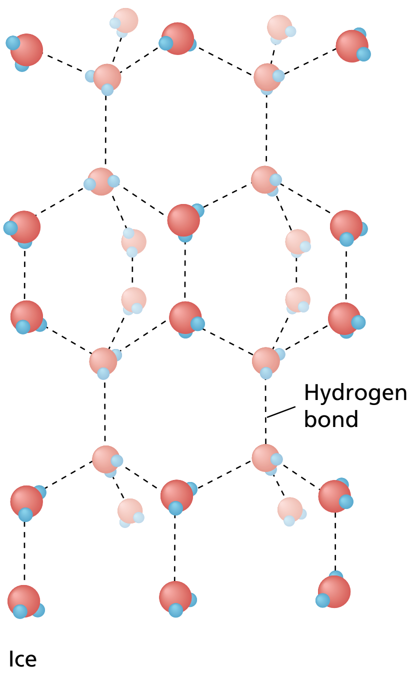


FIGURE 19 Ice contains the same types of bonding as liquid water. However, the structure of the hydrogen bonding is much more rigid and open than it is in liquid water.

molecules is so great that the rigid open structure of the ice crystals breaks down, and ice turns into liquid water.

Figures 18 and 19 also show that the hydrogen bonds between molecules of liquid water at 0.°C are fewer and more disordered than those between molecules of ice at the same temperature. Because the rigid open structure of ice has broken down, water molecules can crowd closer together. Thus, liquid water is denser than ice.

As the liquid water is warmed from 0.°C, the water molecules crowd still closer together. Water molecules are as tightly packed as possible at 3.98°C. At temperatures above 3.98°C, the increasing kinetic energy of the water molecules causes them to overcome molecular attractions. The molecules move farther apart as the temperature continues to rise. As the temperature approaches the boiling point, groups of liquid water molecules absorb enough energy to break up into separate molecules. Because of hydrogen bonding between water molecules, a high kinetic energy is needed, causing water's boiling point to be relatively high (100.°C) compared to other liquids that have similar molar masses.

Physical Properties of Water

At room temperature, pure liquid water is transparent, odorless, tasteless, and almost colorless. Any observable odor or taste is caused by impurities, such as dissolved minerals, liquids, or gases.

As shown by its phase diagram in **Figure 16,** water freezes and ice melts at 0.°C at a pressure of 1 atm (101.3 kPa). The molar enthalpy of fusion of ice is 6.009 kJ/mol. That value is relatively large compared with the molar enthalpy of fusion of other solids. As you have read, water has the unusual property of expanding in volume as it freezes, because its molecules form an open rigid structure. As a result, ice at 0.°C has a density of only about 0.917 g/cm^3, but liquid water at 0.°C has a density of 0.999 84 g/cm^3.

This lower density explains why ice floats in liquid water. The insulating effect of floating ice is particularly important in the case of large bodies of water. If ice were more dense than liquid water, it would sink to the bottom of lakes and ponds, where it would be less likely to melt completely. The water of such bodies of water in temperate climates would eventually freeze solid, killing nearly all the living things in it.

Under a pressure of 1 atm (101.3 kPa), water boils at 100.°C. At this temperature, water's molar enthalpy of vaporization is 40.79 kJ/mol. Both the boiling point and the molar enthalpy of vaporization of water are quite high compared with those of nonpolar substances of comparable molecular mass, such as methane. The values are high because of the strong hydrogen bonding that must be overcome for boiling to occur. The high molar enthalpy of vaporization makes water useful for household steam-heating systems. The steam (vaporized water) stores a great deal of energy as heat. When the steam condenses in radiators, great quantities of energy are released.

SAMPLE PROBLEM A

For more help, go to the Math Tutor at the end of this chapter.

How much energy is absorbed when 47.0 g of ice melts at STP? How much energy is absorbed when this same mass of liquid water boils?

SOLUTION

1 ANALYZE

Given: mass of $H_2O(s)$ = 47.0 g; mass of $H_2O(l)$ = 47.0 g; molar enthalpy of fusion of ice = 6.009 kJ/mol; molar enthalpy of vaporization = 40.79 kJ/mol

Unknown: energy absorbed when ice melts; energy absorbed when liquid water boils

2 PLAN

First, convert the mass of water from grams to moles.

$$47.0\ \cancel{g}\ H_2O \times \frac{1\ \text{mol}\ H_2O}{18.02\ \cancel{g}\ H_2O} = 2.61\ \text{mol}\ H_2O$$

Then, use the molar enthalpy of fusion of a solid to calculate the amount of energy absorbed when the solid melts. Multiply the number of moles by the amount of energy needed to melt one mole of ice at its melting point (the molar enthalpy of fusion of ice). Using the same method, calculate the amount of energy absorbed when water boils by using the molar enthalpy of vaporization.

amount of substance (mol) × molar enthalpy of fusion or vaporization (kJ/mol) = energy (kJ)

3 COMPUTE

$$2.61\ \cancel{\text{mol}} \times 6.009\ \text{kJ}/\cancel{\text{mol}} = 15.7\ \text{kJ (on melting)}$$
$$2.61\ \cancel{\text{mol}} \times 40.79\ \text{kJ}/\cancel{\text{mol}} = 106\ \text{kJ (on vaporizing or boiling)}$$

4 EVALUATE

Units have canceled correctly. The answers have the proper number of significant digits and are reasonably close to estimated values of 18 (3 × 6) and 120 (3 × 40), respectively.

PRACTICE

Answers in Appendix E

1. What quantity of energy is released when 506 g of liquid water freezes?
2. What mass of steam is required to release 4.97×10^5 kJ of energy on condensation?

extension

Go to **go.hrw.com** for more practice problems that ask you to use enthalpies to calculate energies absorbed or released.

Keyword: HC6STMX

SECTION REVIEW

1. Why is a water molecule polar?
2. How is the structure of water responsible for some of water's unique characteristics?
3. Describe the arrangement of molecules in liquid water and in ice.
4. Why does ice float? Why is this phenomenon important?
5. Why is ice less dense than liquid water?
6. Is more energy required to melt one gram of ice at 0°C or to boil one gram of water at 100°C? How do you know?

Critical Thinking

7. **RELATING IDEAS** Why is exposure to steam dangerous?

CHAPTER HIGHLIGHTS

The Kinetic-Molecular Theory of Matter

Vocabulary

kinetic-molecular theory
ideal gas
elastic collision
diffusion
effusion
real gas

- The kinetic-molecular theory of matter can be used to explain the properties of gases, liquids, and solids.
- The kinetic-molecular theory of gases describes a model of an ideal gas.
- Gases consist of large numbers of tiny, fast-moving particles that are far apart relative to their size.

Liquids

Vocabulary

fluid
surface tension
capillary action
vaporization
evaporation
freezing

- The particles of a liquid are closer together and more ordered than those of a gas and are less ordered than those of a solid.
- Liquids have a definite volume and a fairly high density, and they are relatively incompressible. Like gases, liquids can flow and thus are considered to be fluids.

Solids

Vocabulary

crystalline solids
crystal
amorphous solids
melting
melting point
supercooled liquids
crystal structure
unit cell

- The particles of a solid are not nearly as free to move about as those of a liquid or a gas are.
- Solids have a definite shape and may be crystalline or amorphous. They have a definite volume and are generally nonfluid.
- A crystal structure is the total three-dimensional array of points that describes the arrangement of the particles of a crystal.
- Unlike crystalline solids, amorphous solids do not have a highly ordered structure or a regular shape.

Changes of State

Vocabulary

phase
condensation
equilibrium
equilibrium vapor pressure
volatile liquids
boiling
boiling point
molar enthalpy of vaporization
freezing point
molar enthalpy of fusion
sublimation
deposition
phase diagram
triple point
critical point
critical temperature
critical pressure

- A liquid in a closed system will gradually reach a liquid-vapor equilibrium as the rate at which molecules condense equals the rate at which they evaporate.
- When two opposing changes occur at equal rates in the same closed system, the system is said to be in dynamic equilibrium.

Water

- Water is a polar covalent compound.
- The structure and the hydrogen bonding in water are responsible for its relatively high melting point, molar enthalpy of fusion, boiling point, and molar enthalpy of vaporization.

CHAPTER REVIEW

The Kinetic-Molecular Theory of Matter

SECTION 1 REVIEW

1. What idea is the kinetic-molecular theory based on?
2. What is an ideal gas?
3. State the five basic assumptions of the kinetic-molecular theory.
4. How do gases compare with liquids and solids in terms of the distance between their molecules?
5. What is the relationship between the temperature, speed, and kinetic energy of gas molecules?
6. a. What is diffusion?
   b. What factors affect the rate of diffusion of one gas through another?

Liquids

SECTION 2 REVIEW

7. What is a fluid?
8. What is surface tension?
9. Give two reasons why evaporation is a crucial process in nature.

Solids

SECTION 3 REVIEW

10. List six properties of solids, and explain each in terms of the kinetic-molecular theory of solids.
11. List four common examples of amorphous solids.
12. List and describe the four types of crystals in terms of the nature of their component particles and the type of bonding between them.

Changes of State

SECTION 4 REVIEW

13. Using **Figure 14,** estimate the approximate equilibrium vapor pressure of each of the following at the specified temperature.
    a. water at 80°C
    b. diethyl ether at 20°C
    c. ethanol at 60°C
14. a. What is sublimation?
    b. Give two examples of common substances that sublime at ordinary temperatures.
15. What is meant by the normal freezing point of a substance?
16. Explain how the attractive forces between the particles in a liquid are related to the equilibrium vapor pressure of that liquid.
17. Explain the relationship between atmospheric pressure and the actual boiling point of a liquid.
18. Explain the relationship between the molar enthalpy of fusion of a solid and the strength of attraction between that solid's particles.

PRACTICE PROBLEMS

19. a. The molar enthalpy of vaporization for water is 40.79 kJ/mol. Express this enthalpy of vaporization in joules per gram.
    b. The molar enthalpy of fusion for water is 6.009 kJ/mol. Express this enthalpy of fusion in joules per gram.
20. Calculate the molar enthalpy of vaporization of a substance given that 0.433 mol of the substance absorbs 36.5 kJ of energy when it is vaporized.
21. Given that a substance has a molar mass of 259.0 g/mol and a 71.8 g sample of the substance absorbs 4.307 kJ when it melts,
    a. calculate the number of moles in the sample.
    b. calculate the molar enthalpy of fusion.
22. a. Calculate the number of moles in a liquid sample of a substance that has a molar enthalpy of fusion of 3.811 kJ/mol, given that the sample releases 83.2 kJ when it freezes.
    b. Calculate the molar mass of this substance if the mass of the sample is 5519 g.

Water

SECTION 5 REVIEW

23. Describe the structure of a water molecule.

24. List at least eight physical properties of water.

PRACTICE PROBLEMS

25. Which contains more molecules of water: 5.00 cm^3 of ice at 0°C or 5.00 cm^3 of liquid water at 0.°C? How many more? What is the ratio of the numbers of molecules in these two samples?

26. a. What volume and mass of steam at 100.°C and 1.00 atm would release the same amount of energy during condensation as 100. cm^3 of liquid water would release during freezing?

b. What do you note, qualitatively, about the relative volumes and masses of steam and liquid water required to release the same amount of heat? (Hint: See Sample Problem A)

MIXED REVIEW

27. Find the molar enthalpy of vaporization for a substance, given that 3.21 mol of the substance absorbs 28.4 kJ of energy as heat when the substance changes from a liquid to a gas.

28. Water's molar enthalpy of fusion is 6.009 kJ/mol. Calculate the amount of energy as heat required to melt 7.95×10^5 g of ice.

29. A certain substance has a molar enthalpy of vaporization of 31.6 kJ/mol. How much of the substance is in a sample that requires 57.0 kJ to vaporize?

30. Given that water has a molar enthalpy of vaporization of 40.79 kJ/mol, how many grams of water could be vaporized by 0.545 kJ?

31. Calculate the amount of energy released as heat by the freezing of 13.3 g of a liquid substance, given that the substance has a molar mass of 82.9 g/mol and a molar enthalpy of fusion of 4.60 kJ/mol.

32. What volume and mass of steam at 100.°C and 760. torr would release the same amount of energy as heat during condensation as 65.5 cm^3 of liquid water would release during freezing?

33. The following liquid-vapor system is at equilibrium at a given temperature in a closed system.

$$\text{liquid} + \text{energy} \rightleftharpoons \text{vapor}$$

Suppose the temperature is increased and equilibrium is established at the higher temperature. How does the final value of each of the following compare with its initial value? (In each case, answer either higher, lower, or the same.)

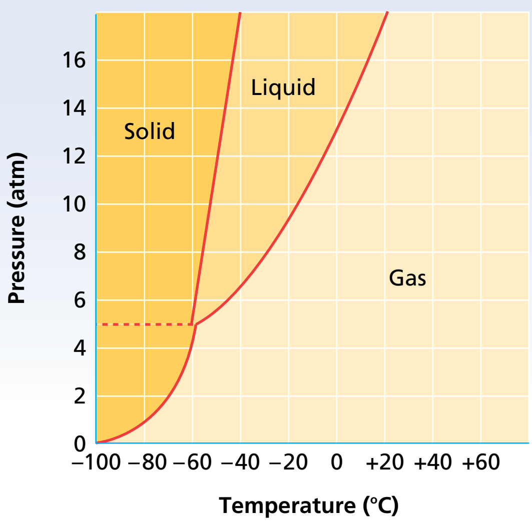
a. the rate of evaporation
b. the rate of condensation
c. the final concentration of vapor molecules
d. the final number of liquid molecules

34. Given a sample of water at any point on curve AB in **Figure 16,** what effect would each of the following changes have on that sample?

a. adding energy at constant pressure
b. decreasing the volume at constant temperature
c. removing energy at constant pressure
d. increasing the volume at constant temperature

35. Using the phase diagram for CO_2, describe all the phase changes that would occur when CO_2 is heated from –100°C to –10°C at a constant pressure of 6 atm.

CO_2 Phase Diagram



CRITICAL THINKING

36. Interpreting Concepts During the freezing of a substance, energy is being removed from that substance. Yet the temperature of the liquid-solid system remains constant. Explain this phenomenon.

37. Applying Models At normal atmospheric pressure, the temperature of an ice-water system remains at 0°C as long as both ice and liquid water are present, regardless of the surrounding temperature. Explain why this occurs.

38. Predicting Outcomes Given a sample of water at any point on curve AD in **Figure 16,** how could more of the liquid water in that sample be converted into a solid without changing the temperature? Explain your reasoning.

39. Interpreting Diagrams Refer to the phase diagram in question 44.

a. Explain what happens when solid CO_2 ("dry ice") warms up to room temperature at normal atmospheric pressure.

b. Is there a pressure below which liquid CO_2 cannot exist? Estimate that pressure from the graph.

USING THE HANDBOOK

40. The *Elements Handbook* contains a table of properties for each group that includes information on the crystal structures of the elements. Most metals crystallize in one of three lattice arrangements: body-centered cubic, face-centered cubic, or hexagonal close-packed. **Figure 10** shows a model of the face-centered cubic lattice for sodium chloride. Use this figure and the information in the *Elements Handbook* to answer the following.

a. What elements in Group 2 have the same lattice structure as sodium chloride?

b. How would the model of an element in a face-centered cubic lattice differ from the compound shown in **Figure 10**?

c. The body-centered cubic lattice is the least-efficient packing structure of the metals. What elements in Groups 1 and 2 show this arrangement?

RESEARCH & WRITING

41. Ceramics are formed from silicates found in the soil. Artists use them to create pottery, but engineers and scientists have created ceramics with superconductive properties. Investigate the growing field of superconductive ceramics.

42. Liquid crystals are substances that possess the combined properties of both liquids and crystals. Write a report on these substances and the various uses we are finding for them.

ALTERNATIVE ASSESSMENT

43. Compile separate lists of crystalline and amorphous solids found in your home. Compare your lists with those of your classmates.

44. Design an experiment to grow crystals of various safe, common household materials. Record the conditions under which each type of crystal is best grown.

extension

Graphing Calculator

Vapor Pressure

Go to **go.hrw.com** for a graphing calculator exercise that asks you to create a graph of the vapor pressure of water as a function of temperature.

Keyword: HC6STMX

Math Tutor CALCULATIONS USING ENTHALPIES OF FUSION

When one mole of a liquid freezes to a solid, energy is released as attractive forces between particles pull the disordered particles of the liquid into a more orderly crystalline solid. When the solid melts to a liquid, the solid must absorb the same quantity of energy in order to separate the particles of the crystal and overcome the attractive forces opposing separation. This quantity of energy used to melt or freeze one mole of a substance at its melting point is called its molar enthalpy of fusion, ΔH_f.

Problem-Solving TIPS

- The enthalpy of fusion of a substance can be given as either joules per gram or kilojoules per mole.
- *Molar* enthalpy of fusion is most commonly used in calculations.
- The enthalpy of fusion is the energy absorbed or given off as heat when a substance melts or freezes at the melting point of the substance.
- No net change in temperature occurs as the state change occurs.

SAMPLE 1

7.30 kJ of energy is required to melt 0.650 mol of ethylene glycol ($C_2H_6O_2$) at its melting point. Calculate the molar enthalpy of fusion, ΔH_f, of ethylene glycol and the energy absorbed.

$$\text{molar enthalpy of fusion} = \Delta H_f = \frac{\text{energy absorbed}}{\text{moles of substance}}$$

$$\Delta H_{f,\text{ethylene glycol}} = \frac{7.30\ \text{kJ}}{0.065\ \text{mol}} = 11.2\ \frac{\text{kJ}}{\text{mol}}$$

SAMPLE 2

Determine the quantity of energy that will be needed to melt 2.50×10^5 kg of iron at its melting point, 1536°C. The ΔH_f of iron is 13.807 kJ/mol.

To calculate the number of moles of iron, use the equation below.

$$\text{moles of substance} = \frac{\text{mass of substance}}{\text{molar mass of substance}}$$

Next, use the following equation for energy as heat absorbed.

$$\text{energy absorbed} = \Delta H_f \times \text{moles of substance}$$

Now, substitute the calculation for moles of substance, and solve.

$$\text{energy absorbed} = \Delta H_f \times \frac{\text{grams of substance}}{\text{molar mass of substance}} =$$

$$13.807\ \frac{\text{kJ}}{\text{mol}} \times \frac{2.50 \times 10^8\ \text{g Fe}}{55.847\ \text{g Fe/mol Fe}}$$

$$\text{energy absorbed} = 6.18 \times 10^7\ \text{kJ}$$

PRACTICE PROBLEMS

1. Calculate the molar enthalpy of fusion of silver if 1.940 mol of silver requires 22.60 kJ of energy to change from a solid to a liquid at its melting point, 961°C.

2. What quantity of energy in kJ must be absorbed by 6.47 mol of solid acetic acid, $C_2H_4O_2$, to melt it at its melting point, 16.7°C? The ΔH_f for acetic acid is 11.54 kJ/mol.

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Surface tension is
A. skin on the surface of a liquid.
B. the tendency of the surface of liquids to decrease the area.
C. the spontaneous mixing of two substances.
D. the same as vapor pressure.

2. Pure liquids boil at higher temperatures under high pressures than they do under low pressures, because
A. the molecules of liquid are closer together under higher pressures.
B. it takes a higher temperature for the vapor pressure to equal the higher external pressure.
C. the molecules of vapor are farther apart under higher pressures.
D. the vapor diffuses more rapidly at higher pressures.

3. The formation of frost is an example of
A. condensation.
B. evaporation.
C. deposition.
D. melting point.

4. The graph that shows the pressure and temperature conditions under which the phases of a substance exist is called
A. a phase diagram.
B. a vapor pressure curve.
C. a unit cell.
D. the kinetic-molecular theory of matter.

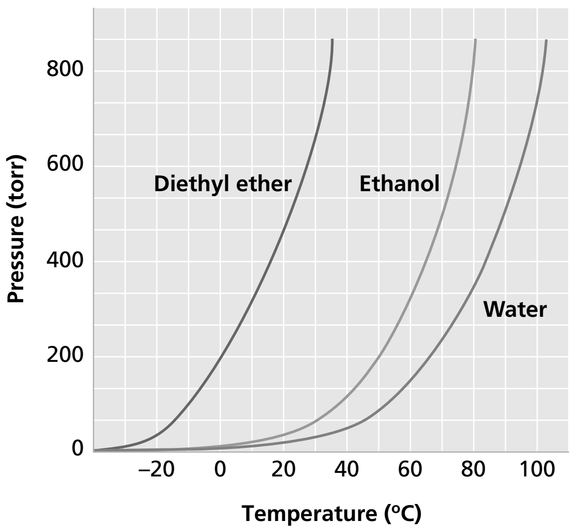
5. Water boils at 100°C. Ethanol boils at 78.5°C. Which of the following statements is true?
A. Water has the higher vapor pressure at 78.5°C.
B. Ethanol has the higher vapor pressure at 78.5°C.
C. Both have the same vapor pressure at 78.5°C.
D. Vapor pressure is not related to boiling point.

6. Which of the following is not a property of typical solids?
A. definite melting point
B. high density
C. easily compressible
D. low rate of diffusion

7. The kinetic-molecular theory states that ideal gas molecules
A. are in constant, rapid, random motion.
B. have mass and take up space.
C. exert forces of attraction and repulsion on each other.
D. have high densities compared with liquids and solids.

SHORT ANSWER

8. Using this graph of vapor pressures of substances at various temperatures, estimate the boiling point of ethanol at an applied (external) pressure of 300 torr.



9. It is found that 60.0 J of energy are required to melt 15 g of a substance. The molar mass of the substance is 120 g/mol. Calculate the enthalpy of fusion of the substance in kilojoules per mole.

EXTENDED RESPONSE

10. Describe how a pressure cooker works.

11. What is meant by the statement that a liquid and its vapor in a closed container are in a state of dynamic equilibrium?

Test TIP Test questions are not necessarily arranged in order of increasing difficulty. If you are unable to answer a question, mark it and move on to other questions.

MICRO-LAB

“Wet” Dry Ice

OBJECTIVES

- *Interpret* a phase diagram.
- *Observe* the melting of CO_2 while varying pressure.
- *Relate* observations of CO_2 to its phase diagram.

MATERIALS

- 4–5 g CO_2 as dry ice, broken into rice-sized pieces
- forceps
- metric ruler
- plastic pipets, 5 mL, shatterproof
- pliers
- scissors
- transparent plastic cup

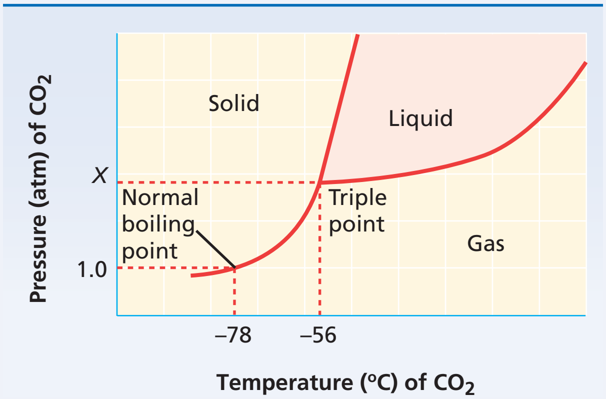


FIGURE A The phase diagram for CO_2 shows the temperatures and pressures at which CO_2 can undergo phase changes.

BACKGROUND

The phase diagram for carbon dioxide in **Figure A** shows that CO_2 can exist only as a gas at ordinary room temperature and pressure. To observe the transition of solid CO_2 to liquid CO_2, you must increase the pressure until it is at or above the triple point pressure, which is labeled X in the diagram.

SAFETY

For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

PREPARATION

1. Organize a place in your lab notebook for recording your observations.

PROCEDURE

1. Use forceps to place 2–3 very small pieces of dry ice on the table, and observe them until they have completely sublimed. **Caution:** Dry ice will freeze skin very quickly. Do not attempt to pick up the dry ice with your fingers.

2. Fill a plastic cup with tap water to a depth of 4–5 cm.

3. Cut the tapered end (tip) off the graduated pipet.

4. Use forceps to carefully slide 8–10 pieces of dry ice down the stem and into the bulb of the pipet.

5. Use a pair of pliers to clamp the opening of the pipet stem securely shut so that no gas can escape. Use the pliers to hold the tube and to lower the pipet into the cup just until the bulb is submerged, as shown in **Figure B.** From the side of the cup, observe the behavior of the dry ice.

6. As soon as the dry ice has begun to melt, quickly loosen the pliers while still holding the bulb in the water. Observe the CO_2.

7. Tighten the pliers again, and observe.

8. Repeat Procedure steps 6 and 7 as many times as possible.

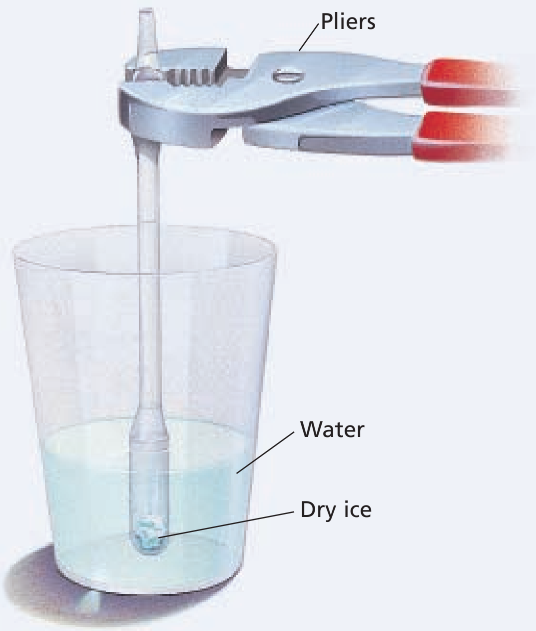


FIGURE B Clamp the end of the pipet shut with the pliers. Submerge the bulb in water in a transparent cup.

CLEANUP AND DISPOSAL

9. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or place them in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. **Analyzing Results:** What differences did you observe between the subliming and the melting of CO_2?

2. **Analyzing Methods:** As you melted the CO_2 sample over and over, why did it eventually disappear? What could you have done to make the sample last longer?

3. **Analyzing Methods:** What purpose(s) do you suppose the water in the cup served?

EXTENSIONS

1. **Predicting Outcomes:** What would have happened if fewer pieces of dry ice (only 1 or 2) had been placed inside the pipet bulb? If time permits, test your prediction.

2. **Predicting Outcomes:** What might have happened if too much dry ice (20 or 30 pieces, for example) had been placed inside the pipet bulb? How quickly would the process have occurred? If time permits, test your prediction.

3. **Predicting Outcomes:** What would have happened if the pliers had not been released once the dry ice melted? If time permits, test your prediction.