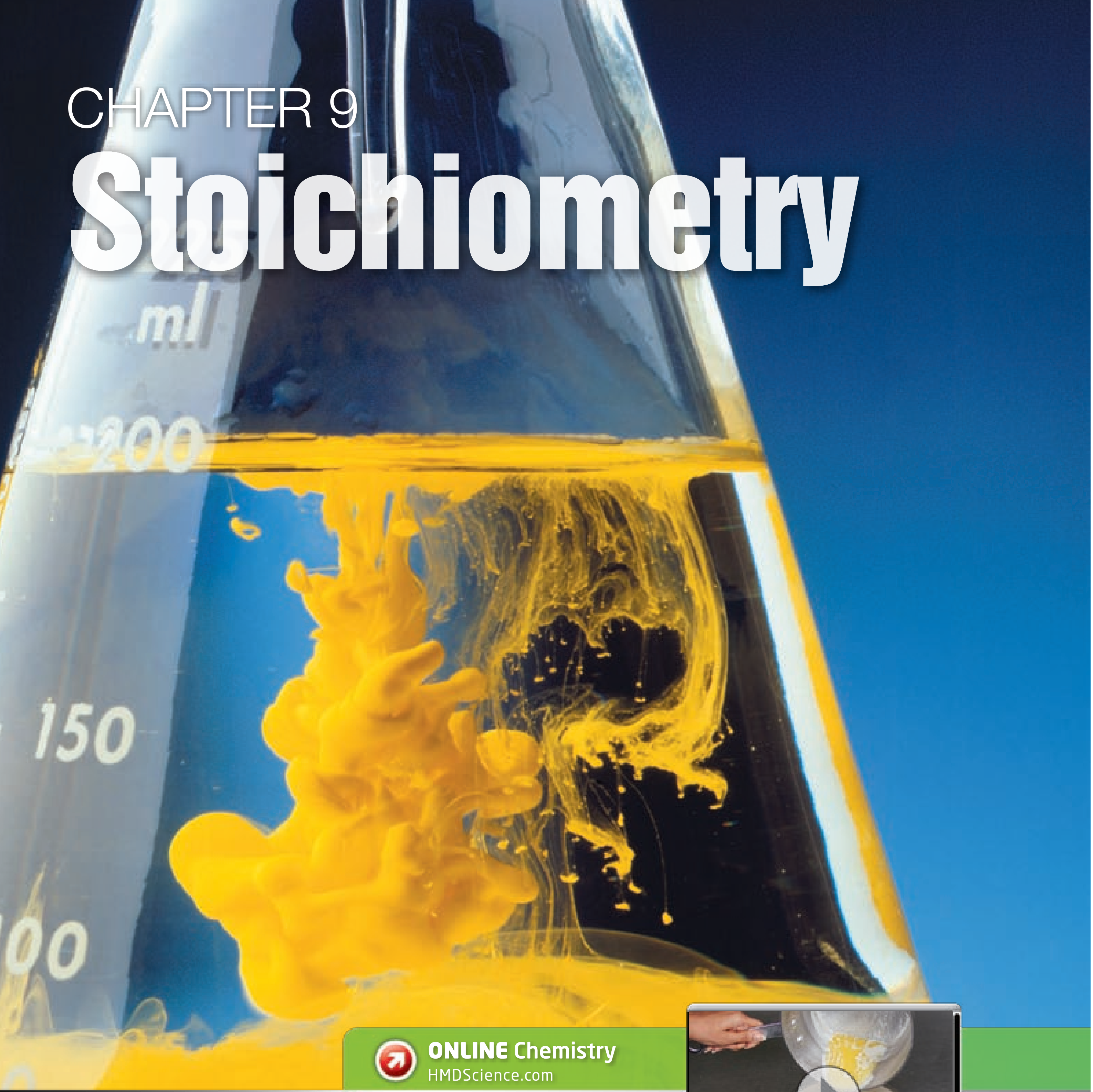
CHAPTER 9

Stoichiometry



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SECTION 1
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Stoichiometry

SECTION 1

Main Idea

Ratios of substances in chemical reactions can be used as conversion factors.

Introduction to Stoichiometry

Key Terms

composition stoichiometry
reaction stoichiometry
mole ratio

VIRGINIA STANDARDS

CH.4 The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include:
CH.4.a Avogadro's principle and molar volume.
CH.4.b stoichiometric relationships.
CH.1.EKS-19; **CH.1.EKS-26**; **CH.4.EKS-2**

Much of our knowledge of chemistry is based on the careful quantitative analysis of substances involved in chemical reactions. **Composition stoichiometry 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, the subject of this chapter, 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.

MAIN IDEA

Ratios of substances in chemical reactions can be used as conversion factors.

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.

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

mass of *given* substance (g)	$\longrightarrow$	amount of *given* substance (mol)	$\longrightarrow$	amount of *unknown* substance (mol)

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

mass of *given* substance (g)	$\longrightarrow$	amount of *given* substance (mol)	$\longrightarrow$	amount of *unknown* substance (mol)	$\longrightarrow$	mass of *unknown* substance (g)

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 your study of equations and reactions 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.

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

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

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

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

CAREERS IN CHEMISTRY

Chemical Technician S.T.E.M.

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.

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.

Molar Mass

Recall 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.

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

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

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

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 1 FORMATIVE ASSESSMENT

Reviewing Main Ideas

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.

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.

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* (Greek, meaning "burned"), 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.

Antoine Laurent Lavoisier helped establish chemistry as a science.

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.

TABLE OF SIMPLE SUBSTANCES

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

New Names	*Correspondent old Names*
Light - - -	*Light.*
Caloric - - -	*Heat. Principle or element of heat. Fire. Igneous fluid. Matter of fire and of heat.*
Oxygen - - -	*Dephlogisticated air. Empyreal air. Vital air, or Base of vital air.*
Azote - - -	*Phlogisticated air or gas. Mephitis, or its base.*
Hydrogen - -	*Inflammable air or gas, or the base of inflammable air.*

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

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"?

SECTION 2

Main Ideas

- Balanced equations give amounts of reactants and products under ideal conditions.
- Mole-to-gram calculations require two conversion factors.
- Gram-to-mole conversions require the molar mass of the given substance and the mole ratio.
- Mass-to-mass calculations use the mole ratio and the molar masses of the given and unknown substances.

VIRGINIA STANDARDS

CH.4.b The student will investigate and understand that chemical quantities are based on molar relationships. Key concepts include: stoichiometric relationships.
CH.1.EKS-19; **CH.4.EKS-2**

Ideal Stoichiometric Calculations

A 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 reaction 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 when by-products are formed in addition to the expected products.

Solving stoichiometric problems requires practice. Practice by working the sample problems in the rest of this chapter. Using a logical, systematic approach will help you successfully solve these problems.

MAIN IDEA

Balanced equations give amounts of reactants and products under ideal conditions.

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

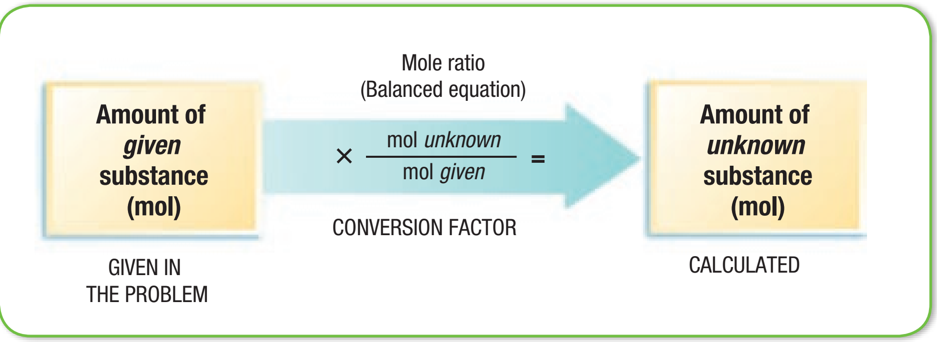
$$\begin{matrix}\text{amount of} \\ \textit{given}\text{ substance (mol)}\end{matrix} \longrightarrow \begin{matrix}\text{amount of} \\ \textit{unknown}\text{ substance (mol)}\end{matrix}$$

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 (see **Figure 2.1**).

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

FIGURE 2.1

Mole-to-Mole Conversions This is a solution plan for problems in which the given and unknown quantities are expressed in moles.



Stoichiometric Calculations Using Mole Ratios

Sample Problem A **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.**

$$\mathbf{CO_2(\mathit{g}) + 2LiOH(\mathit{s}) \longrightarrow Li_2CO_3(\mathit{s}) + H_2O(\mathit{l})}$$

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

1 ANALYZE

Given: amount of CO_2 = 20 mol

Unknown: amount of LiOH (mol)

2 PLAN

$$\text{amount of } CO_2 \text{ (mol)} \longrightarrow \text{amount of LiOH (mol)}$$

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 SOLVE

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 CHECK YOUR WORK

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 as an ingredient 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?

MAIN IDEA

Mole-to-gram calculations require two conversion factors.

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 (see **Figure 2.2** on the next page).

CHECK FOR UNDERSTANDING

Explain Why might it be important for a chemist to be able to calculate, in advance, how much product a chemical reaction would yield?

Stoichiometric Calculations Using Mole Ratios

Sample Problem B **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?**

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 SOLVE

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

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

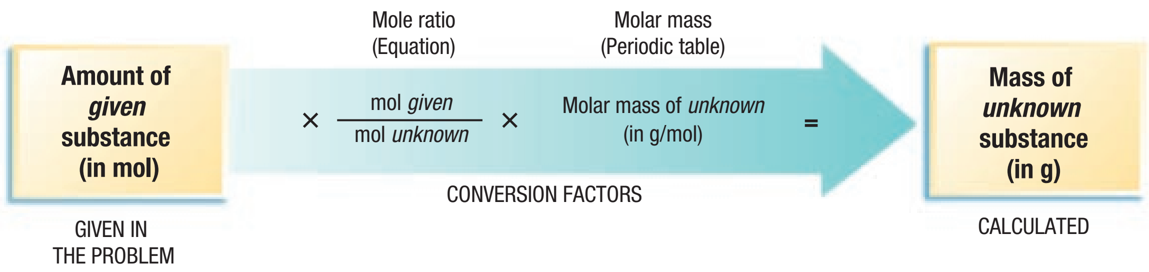
$$3.00 \cancel{\text{mol } H_2O} \times \frac{1 \cancel{\text{mol } C_6H_{12}O_6}}{6 \cancel{\text{mol } H_2O}} \times \frac{180.18 \text{ g } C_6H_{12}O_6}{1 \cancel{\text{mol } C_6H_{12}O_6}} = 90.1 \text{ g } C_6H_{12}O_6$$

4 CHECK YOUR WORK

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.

FIGURE 2.2

Mole-to-Gram Conversions This is a solution plan for problems in which the given quantity is expressed in moles, and the unknown quantity is expressed in grams.



Stoichiometric Calculations Using Mole Ratios

Sample Problem C **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?**

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.

3 SOLVE

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

Use the periodic table to compute the molar mass of CO_2.

$CO_2 = 44.01$ g/mol

4 CHECK YOUR WORK

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

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?

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

MAIN IDEA

Gram-to-mole conversions require the molar mass of the given substance and the mole ratio.

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

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

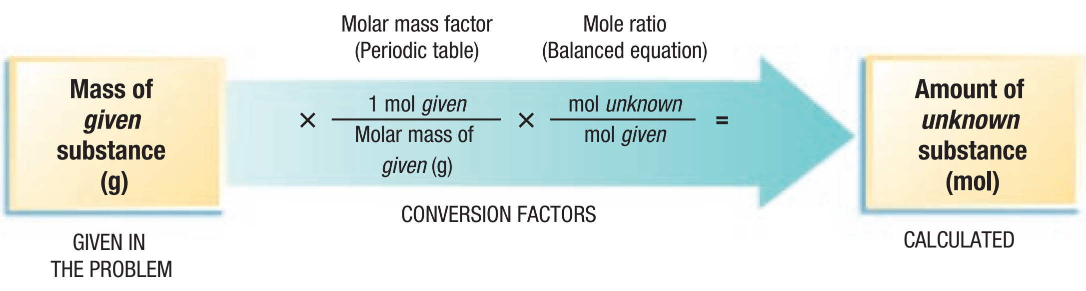
This route requires two additional pieces of data: the molar mass of the *given* substance and the mole ratio. The molar mass is determined by using masses from the periodic table. We will follow a procedure much like the one used previously by using the units of the molar mass conversion factor to guide 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 (see **Figure 2.3** below).

You should take time at this point to look at each type of stoichiometric conversion and make note that in every type, you must begin with a correctly balanced chemical equation. It is important to remember that without a balanced equation, you will not have an accurate molar ratio and will not be able to calculate the correct molar mass to use in your conversions.

If you are still struggling with balancing equations, it is highly recommended that you continue to practice the process. The skill of balancing chemical equations will be used continuously throughout the remainder of the course.

FIGURE 2.3

Gram-to-Mole Conversions This is a solution plan for problems in which the given quantity is expressed in grams, and the unknown quantity is expressed in moles.



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Stoichiometric Calculations Using Mole Ratios

Sample Problem D **The first step in the industrial manufacture of nitric acid is the catalytic oxidation of ammonia.**

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

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?

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.

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

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 SOLVE

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

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

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

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 CHECK YOUR WORK

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?

FIGURE 2.4

Mass-to-Mass Conversions 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.

Molar mass factor (Periodic table) | Mole ratio (Balanced equation) | Molar mass factor (Periodic table)

$$\text{Mass of } \textit{given} \text{ substance (g)} \times \frac{1 \text{ mol } \textit{given}}{\text{Molar mass of } \textit{given} \text{ (g)}} \times \frac{\text{mol } \textit{unknown}}{\text{mol } \textit{given}} \times \frac{\text{Molar mass of } \textit{unknown} \text{ (g)}}{1 \text{ mol } \textit{unknown}} = \text{Mass of } \textit{unknown} \text{ substance (g)}$$

GIVEN IN THE PROBLEM | CONVERSION FACTORS | CALCULATED

MAIN IDEA

Mass-to-mass calculations use the mole ratio and the molar masses of the given and unknown substances.

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

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

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 (see **Figure 2.4** above).

Stoichiometric Calculations Using Mole Ratios

Sample Problem E **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.**

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

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

1 ANALYZE

Given: amount of HF = 3.00 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.

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

Continued

3 SOLVE

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

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

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

4 CHECK YOUR WORK

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?

SECTION 2 FORMATIVE ASSESSMENT

Reviewing Main Ideas

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

Main Ideas

- One reactant limits the product of a reaction.
- Comparing the actual and theoretical yields helps chemists determine the reaction's efficiency.

Limiting Reactants and Percentage Yield

Key Terms

limiting reactant
excess reactant
theoretical yield
actual yield
percentage yield

VIRGINIA STANDARDS

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include:
CH.1.a designated laboratory techniques.
CH.1.b safe use of chemicals and equipment.
CH.4.b The student will investigate and understand that chemical quantities are based on molar relationships. Key concepts include: stoichiometric relationships.
CH.4.EKS-3; CH.4.EKS-4

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.

MAIN IDEA

One reactant limits the product of a reaction.

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.*

Consider the reaction between carbon and oxygen:

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

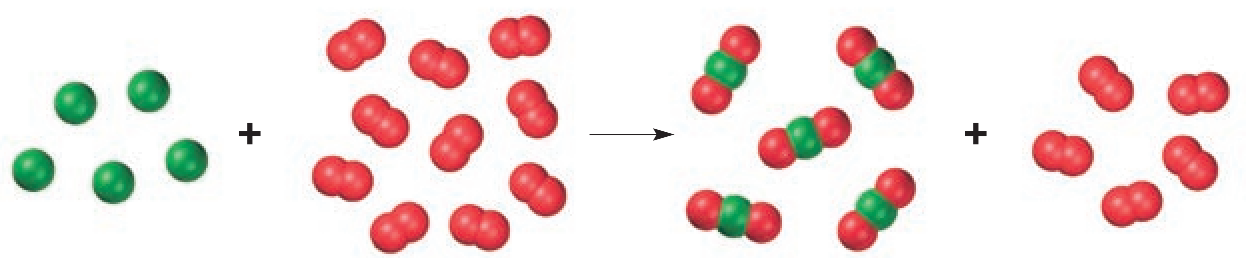
According to the equation, one mole of carbon reacts with one mole of oxygen to form one mole of carbon dioxide. Suppose you could mix 5 mol C with 10 mol O_2 and allow the reaction to take place. **Figure 3.1** shows that there is more oxygen than is needed to react with the carbon. Carbon is the limiting 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 3.1

Limiting Reactants Carbon and oxygen always react in the same molar ratio to form carbon dioxide.

CRITICAL THINKING

Explain Which substance is the limiting reactant in this chemical reaction? Explain your answer.



PREMIUM CONTENT

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Limiting Reactant

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 are added to 4.5 mol SiO_2, which is the limiting reactant?

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 SOLVE

$$6.0 \cancel{\text{mol HF}} \times \frac{1 \text{ mol } SiF_4}{4 \cancel{\text{mol HF}}} = 1.5 \text{ mol } SiF_4 \text{ produced}$$

$$4.5 \cancel{\text{mol } SiO_2} \times \frac{1 \text{ mol } SiF_4}{1 \cancel{\text{mol } SiO_2}} = 4.5 \text{ mol } SiF_4 \text{ produced}$$

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 CHECK YOUR WORK

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?

Limiting Reactant

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 are 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?

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.

$$\text{original g excess reactant} - \text{g excess reactant consumed} = \text{g excess reactant remaining}$$

Continued

3 SOLVE

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

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

$36.0\ \text{g H}_2\text{O} - 28.8\ \text{g H}_2\text{O consumed} = 7.2\ \text{g H}_2\text{O remaining}$

4 CHECK YOUR WORK

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?

Quick LAB — 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 un-greased 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?

MAIN IDEA

Comparing the actual and theoretical yields helps chemists determine the reaction's efficiency.

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 by-products 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.**

Percentage Yield $= \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$

CHECK FOR UNDERSTANDING

Explain Explain why it might be necessary for a chemist in industry to find out how efficient a chemical reaction is.

Percentage Yield

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?

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.

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

Continued

3 SOLVE

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

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

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

(theoretical yield)

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

4 CHECK YOUR WORK

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 react 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?

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

SECTION 3 FORMATIVE ASSESSMENT

Reviewing Main Ideas

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 than, the same as, or less than 15 mol? How many moles of H_2 are needed?

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 Problem

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

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.

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

CHAPTER 9 Summary

PREMIUM CONTENT
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SECTION 1 Introduction to Stoichiometry

- 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.

KEY TERMS

composition stoichiometry
reaction stoichiometry
mole ratio

SECTION 2 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.

SECTION 3 Limiting Reactants and 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.

KEY TERMS

limiting reactant
excess reactant
theoretical yield
actual yield
percentage yield

CHAPTER 9 Review

SECTION 1

Introduction to Stoichiometry

REVIEWING MAIN IDEAS

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.

SECTION 2

Ideal Stoichiometric Calculations

REVIEWING MAIN IDEAS

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. During lightning flashes, nitrogen combines with oxygen in the atmosphere 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 Problem 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.

SECTION 3

Limiting Reactants and Percentage Yield

REVIEWING MAIN IDEAS

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 5.00kg of ammonia?

Mixed Review

REVIEWING MAIN IDEAS

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 of 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, $CaSO_4 \bullet \frac{1}{2}H_2O$, in airtight containers to prevent it from absorbing water vapor from the air and changing to gypsum, $CaSO_4 \bullet 2H_2O$. 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* (Appendix A) 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* (Appendix A). 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* (Appendix A). 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 AND 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.

Standards-Based Assessment

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 are 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.