CHAPTER 9

Stoichiometry

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

Introduction to Stoichiometry

SECTION 9-1

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 matter. All reaction-stoichiometry calculations start with a balanced chemical equation. This equation gives the relative numbers of moles of reactants and products.

CHEMISTRY INTERACTIVE TUTOR
Module 5: Equations and Stoichiometry

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

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

Problem Type 2: *Given* is an amount in moles and the *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

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

Problem Type 3: *Given* is a mass in grams and the *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

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

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

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

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 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 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.

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.

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

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

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

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

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

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

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

SECTION REVIEW

1. What is stoichiometry?
2. How is a mole ratio from a reaction used in stoichiometric problems?
3. For each of the following chemical equations, write all possible mole ratios.
   a. $2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$
   b. $4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(l)$
   c. $2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

The Case of Combustion

HISTORICAL PERSPECTIVE

People throughout history have transformed substances by burning them in the 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 eighteenth 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 seventeenth 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. Though early chemists experimented extensively, most disregarded the importance of measurement, an oversight that set chemistry on the wrong path 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 heat, smoke, and light. To account for it, a theory was proposed that combustion depended on the emission of a substance called phlogiston, which appeared as a combination of heat and light while the material was burning but which couldn't be detected beforehand.

Antoine-Laurent Lavoisier and his wife, Marie-Anne Pierrette Lavoisier, who assisted him. One of her important roles was to translate the papers of important scientists for her husband.

The Metropolitan Museum of Art, Purchase, Mr. and Mrs. Charles Wrightsman Gift, in honor of Everett Fahy, 1977. (1977.10) Copyright © 1989 By The Metropolitan Museum of Art.

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 he had obtained "phlogisticated air."

A couple of years later, Joseph Priestley found that when he heated mercury in air, he obtained a reddish powder, which he assumed to be mercury devoid of phlogiston. But when he decided to heat the powder, he recorded an unexpected result:

> *I endeavored to extract air from [the powder by heating it]; and I presently found that*

... air was expelled from it readily. Having got about three or four times as much as the bulk of my materials, I admitted water to it, and found that water was not imbibed by it. But what surprised me more ... was, that a candle in this air burned ... remarkably ...

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 weighed anything. Unlike his colleagues, Lavoisier knew the importance of using a balance:

... making experiments ... is founded on this principle ... always suppose an exact equality or equation between the principles [masses] of the body examined and those of the products of its analysis.

Applying this rule, which would become known as the law of conservation of mass, Lavoisier endeavored to explain the results of Rutherford and Priestley.

He 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 during combustion. He then weighed the burnt metal and observed a weight increase relative to the original tin. Curiously, this increase equaled the weight of the air that had rushed into the vessel. To Lavoisier, this did not support the idea of phlogiston escaping the burning material. Instead, it indicated that during combustion a portion of air was depleted.

After obtaining similar results using a variety of substances, Lavoisier concluded that air was not an element at all 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, with 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 theories of phlogiston and 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, earning him the reputation as the father of chemistry.

internet connect

SCI LINKS NSTA

TOPIC: Lavoisier
GO TO: www.scilinks.org
***sci*LINKS CODE:** HC2091

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 as published in his book Elements of Chemistry *in 1789.*

SECTION 9-2

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 chemical equation plays a very important part 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 make predictions about chemical reactions 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, ideal conditions are rarely met in the laboratory or in industry. Yet, theoretical stoichiometric calculations serve the very important function of showing the maximum amount of product that could be obtained before a reaction is run in the laboratory.

Solving stoichiometric problems requires practice. These problems are extensions of the composition-stoichiometry problems 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

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

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.

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

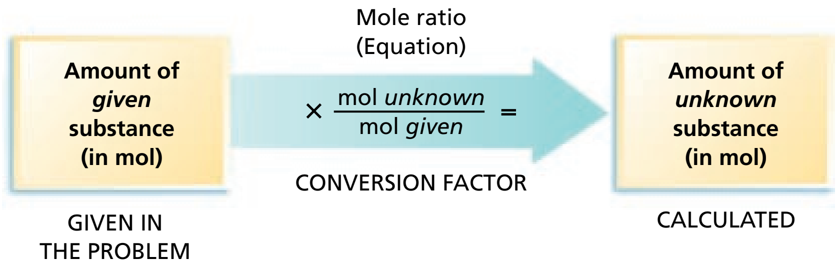


FIGURE 9-1 This is a solution plan for problems in which the given and unknown quantities are expressed in moles.

SAMPLE PROBLEM 9-1

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 of 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 in moles

2 PLAN

amount of CO_2 (in mol) $\longrightarrow$ amount of LiOH (in 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 give you mol LiOH in your final answer. The correct ratio is the following.

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

This ratio 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 rounded correctly to one significant figure to match that in the factor 20 mol CO_2, and the units cancel to leave mol LiOH, which is the unknown. The equation shows that twice the amount in moles of LiOH react with CO_2. Therefore, the answer should be greater than 20.

PRACTICE

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? *Answer* 4 mol NH_3

2. The decomposition of potassium chlorate, $KClO_3$, is used as a source of oxygen in the laboratory. How many moles of potassium chlorate are needed to produce 15 mol of oxygen? *Answer* 10. mol $KClO_3$

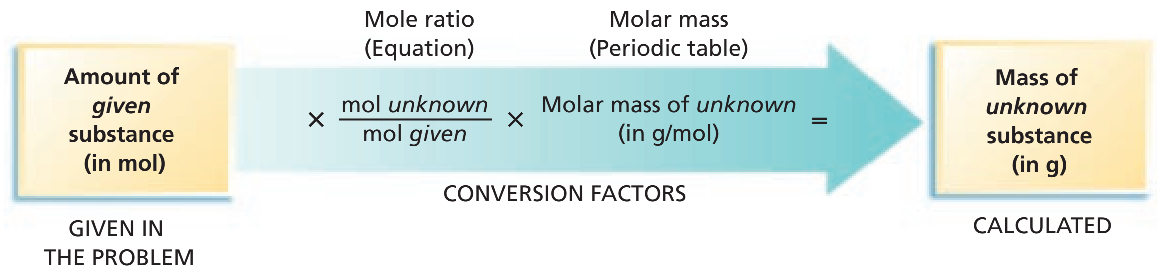
Conversions of Amounts in Moles to Mass

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

amount of *given* substance (in mol) ⟶ amount of *unknown* substance (in mol) ⟶ mass of *unknown* substance (in g)

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 9-2 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 9-2

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 (in 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 $C_6H_{12}O_6$.

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

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 9-3

What mass of carbon dioxide, in grams, is needed to react with 3.00 mol of H_2O in the photosynthetic reaction described in Sample Problem 9-2?

SOLUTION

1 *ANALYZE*

Given: amount of H_2O = 3.00 mol

Unknown: mass of CO_2 in grams

2 *PLAN* The chemical equation from Sample Problem 9-2 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 of CO_2.

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

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

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

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

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

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?

Answer
80.6 g MgO

2. What mass in grams of oxygen combines with 2.00 mol of magnesium in this same reaction?

Answer
32.0 g O_2

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

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

Answer
300 g $C_6H_{12}O_6$

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

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

This route also requires two additional pieces of data: the molar mass of the *given* substance and the mole ratio. The molar mass is determined using masses from the periodic table. To convert the mass of a substance to moles we are using a factor which we will call the inverted molar mass. It is simply one over the molar mass. To solve this type of problem, simply multiply or divide the known quantity by the appropriate conversion factors as follows.

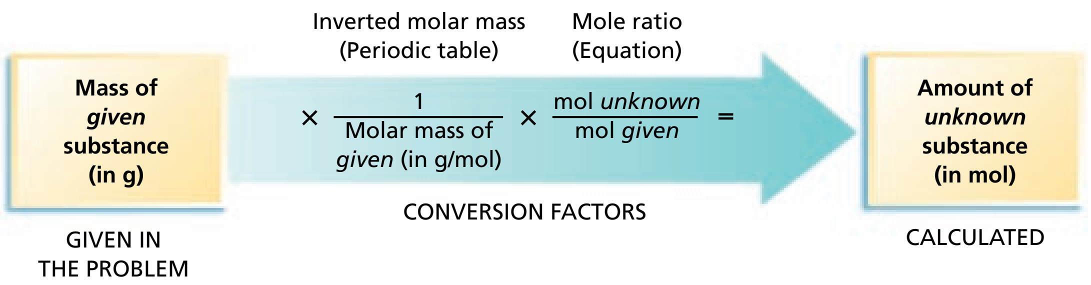


FIGURE 9-3 This is a solution plan for problems in which the given quantity is expressed in grams and the unknown quantity is expressed in moles.

SAMPLE PROBLEM 9-4

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 of 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 (in mol)
b. amount of H_2O produced (in 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 of 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 inverted molar mass of NH_3.

$$\textbf{a.}\ \text{g } NH_3 \times \overbrace{\frac{1 \text{ mol } NH_3}{\text{g } NH_3}}^{\text{inverted molar mass}} \times \overbrace{\frac{\text{mol NO}}{\text{mol } NH_3}}^{\text{mol ratio}} = \text{mol NO}$$

$$\textbf{b.}\ \text{g } NH_3 \times \overbrace{\frac{1 \text{ mol } NH_3}{\text{g } NH_3}}^{\text{inverted molar mass}} \times \overbrace{\frac{\text{mol } H_2O}{\text{mol } NH_3}}^{\text{mol ratio}} = \text{mol } H_2O$$

3 COMPUTE

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

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

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

$$\textbf{b.}\ 824 \cancel{\text{g } NH_3} \times \frac{1 \cancel{\text{mol } NH_3}}{17.04 \cancel{\text{g } NH_3}} \times \frac{6 \text{ mol } H_2O}{4 \cancel{\text{mol } NH_3}} = 72.6 \text{ mol } H_2O$$

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

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? *Answer* 7.81 mol HgO

2. How many moles of mercury are produced? *Answer* 7.81 mol Hg

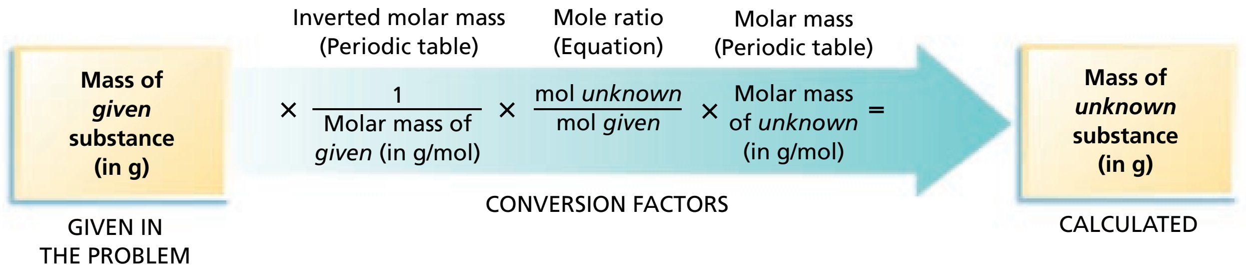


FIGURE 9-4 This is a solution plan for problems in which the given quantity is expressed in grams and the unknown quantity is also expressed in grams.

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

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

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 9-5

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.

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

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

SOLUTION

1 ANALYZE

Given: amount of HF = 30.00 g
Unknown: mass of SnF_2 produced in grams

2 PLAN

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

$$\text{g HF} \times \overbrace{\frac{\text{1 mol HF}}{\text{g HF}}}^{\text{inverted molar mass}} \times \overbrace{\frac{\text{mol SnF}_2}{\text{mol HF}}}^{\text{mol ratio}} \times \overbrace{\frac{\text{g SnF}_2}{\text{mol SnF}_2}}^{\text{molar mass}} = \text{g SnF}_2$$

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

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

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

4 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

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 of 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 of Cu?

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

Answer

1. a. 60.0 g NH_4NO_3

b. 27.0 g H_2O

2. 339 g

3. 2.6 kg

SECTION REVIEW

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

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

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 mol 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 of CaC_2 are consumed in this reaction, how many moles of H_2O are needed?
b. How many moles of each product would be formed?

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

5. Acetylene gas, C_2H_2, used in welding, produces an extremely hot flame when it burns in pure oxygen according to the following reaction.

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g)$$

How many grams of each product are produced when 2.50×10^4 g of C_2H_2 burns completely?

SECTION 9-3

Limiting Reactants and Percent 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 percent yield.
- Calculate percent yield, given the actual yield and quantity of a reactant.

In the laboratory, a reaction is rarely carried out with exactly the required amounts of each of the reactants. In most 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 amounts of the other reactants 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 sometimes called the* **excess reactant.** A limiting reactant may also be referred to as a limiting reagent.

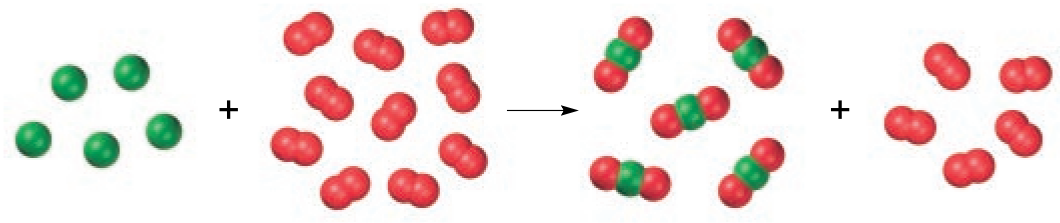
The concept of the limiting reactant is analogous to the relationship between the number of people who want to take a certain airplane flight and the number of seats available in the airplane. If 400 people want to travel on the flight and only 350 seats are available, then only 350 people can go on the flight. The number of seats on the airplane limits the number of people who can travel. There are 50 people in excess.

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 of C with 10 mol of O_2 and allow the reaction to take place. Figure 9-5 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 of O_2 will be left over at the end of the reaction.

FIGURE 9-5 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 9-6

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 2.0 mol of HF are exposed to 4.5 mol of SiO_2, which is the limiting reactant?

SOLUTION

1 ANALYZE

Given: amount of HF = 2.0 mol
amount of SiO_2 = 4.5 mol
Unknown: limiting reactant

2 PLAN

The given amount of either reactant is used to calculate the required amount of the other reactant. The calculated amount is then compared with the amount actually available, and the limiting reactant can be identified. We will choose to calculate the moles of SiO_2 required by the given amount of HF.

$$\text{mol HF} \times \frac{\text{mol } SiO_2}{\text{mol HF}} = \text{mol } SiO_2 \text{ required}$$

3 COMPUTE

$$2.0 \cancel{\text{mol HF}} \times \frac{1 \text{ mol } SiO_2}{4 \cancel{\text{mol HF}}} = 0.50 \text{ mol } SiO_2 \text{ required}$$

Under ideal conditions, the 2.0 mol of HF will require 0.50 mol of SiO_2 for complete reaction. Because the amount of SiO_2 available (4.5 mol) is more than the amount required (0.50 mol), the limiting reactant is HF.

4 EVALUATE

The calculated amount of SiO_2 is correctly given to two significant figures. Because each mole of SiO_2 requires 4 mol of HF, it is reasonable that HF is the limiting reactant because the molar amount of HF available is less than half that of SiO_2.

PRACTICE

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 of N_2H_4 is mixed with 0.500 mol of H_2O_2?

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

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

2. If 20.5 g of chlorine is reacted with 20.5 g of sodium, which reactant is in excess? How do you know?

Answer

1. a. H_2O_2

b. 0.500 mol N_2H_4

c. 0.250 mol N_2, 1.00 mol H_2O

2. Sodium is in excess because only 0.578 mol Na is needed.

SAMPLE PROBLEM 9-7

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 of H_2O is mixed with 167 g of 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 = 167 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. Choose one reactant and calculate the needed amount of the other to determine which is the limiting reactant. We have chosen Fe. The mole ratio from the balanced equation is 3 mol Fe for every 4 mol H_2O.

$$\text{g Fe} \times \overset{\text{inverted molar mass}}{\frac{1 \text{ mol Fe}}{\text{g Fe}}} = \text{mol Fe available}$$

$$\text{g } H_2O \times \overset{\text{inverted molar mass}}{\frac{1 \text{ mol } H_2O}{\text{g } H_2O}} = \text{mol } H_2O \text{ available}$$

$$\text{mol Fe} \times \overset{\text{mol ratio}}{\frac{\text{mol } H_2O}{\text{mol Fe}}} = \text{mol } H_2O \text{ required}$$

b. To find the maximum amount of Fe_3O_4 that can be produced, the given amount in moles of the limiting reactant must be used in a simple stoichiometric problem.

$$\text{limiting reactant (in mol)} \times \overset{\text{mol ratio}}{\frac{\text{mol } Fe_3O_4}{\text{mol limiting reactant}}} \times \overset{\text{molar mass}}{\frac{\text{g } Fe_3O_4}{\text{mol } Fe_3O_4}} = \text{g } Fe_3O_4 \text{ produced}$$

c. To find the amount of excess reactant remaining, the amount of the excess reactant that is consumed must first be determined. The given amount in moles of the limiting reactant must be used in a simple stoichiometric problem.

$$\begin{matrix}\text{limiting reactant}\\ \text{(in mol)}\end{matrix} \times \frac{\text{mol excess reactant}}{\text{mol limiting reactant}} \times \frac{\text{g excess reactant}}{\text{mol excess reactant}} = \begin{matrix}\text{g excess reactant}\\ \text{consumed}\end{matrix}$$

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 Use the periodic table to determine the molar masses of H_2O, Fe, and Fe_3O_4.

H_2O = 18.02 g/mol

Fe = 55.85 g/mol

Fe_3O_4 = 231.55 g/mol

$$36.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 2.00 \text{ mol } H_2O$$

$$167 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 2.99 \text{ mol Fe}$$

$$2.99 \cancel{\text{ mol Fe}} \times \frac{4 \text{ mol } H_2O}{3 \cancel{\text{ mol Fe}}} = 3.99 \text{ mol } H_2O \text{ required}$$

a. The required 3.99 mol of H_2O is more than the 2.00 mol of H_2O available, so H_2O is the limiting reactant.

b. $$2.00 \text{ mol } H_2O \times \frac{1 \text{ mol } Fe_3O_4}{4 \text{ mol } H_2O} \times \frac{231.55 \text{ g } Fe_3O_4}{\text{mol } Fe_3O_4} = 116 \text{ g } Fe_3O_4$$

c. $$2.00 \text{ mol } H_2O \times \frac{3 \text{ mol Fe}}{4 \text{ mol } H_2O} \times \frac{55.85 \text{ g Fe}}{\text{mol Fe}} = 83.8 \text{ g Fe consumed}$$

167 g Fe originally present – 83.8 g Fe consumed = 83.2 g Fe remaining

4 EVALUATE Three significant digits are carried through each calculation. The result of the final subtraction is rounded to match the significance of the least accurately known number, that is, the units digit for the original mass of Fe. The mass of Fe_3O_4 is close to an estimated answer of 115, which is one-half of 230. The amount of the limiting reactant, H_2O, is about one-half the amount needed to use all of the Fe, so about one-half the Fe remains unreacted.

PRACTICE

CHEMISTRY INTERACTIVE TUTOR
Module 5: Equations and Stoichiometry

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

$$8Zn(s) + S_8(s) \longrightarrow 8ZnS(s)$$

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?

Answer

1. a. Zn

b. 0.75 mol S_8 remains

c. 2.00 mol ZnS

2. a. carbon

b. 2.40 mol H_2 and 2.40 mol CO

c. 4.85 g H_2 and 67.2 g CO

QuickLAB

Limiting Reactants in a Recipe

Wear oven mitts when handling heated items.

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 ingredients 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?

Percent Yield

The amounts of products calculated in the stoichiometric 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. Some of the reactant may be used in competing side reactions that reduce the amount of the desired product. Also, once a product is formed, it often is usually collected in impure form, and some of the product is often lost during the purification process. *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* **percent yield** *is the ratio of the actual yield to the theoretical yield, multiplied by 100.*

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

SAMPLE PROBLEM 9-8

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.

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

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

2 *PLAN*

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

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

Then the percent yield can be found.

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

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

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

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

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

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 percent yield is close to an estimated value of 80%, ($40/50 \times 100$).

PRACTICE

1. Methanol can be produced through the reaction of CO and H_2 in the presence of a catalyst. *Answer* 79.8%

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

If 75.0 g of CO reacts to produce 68.4 g CH_3OH, what is the percent 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 percent yield of Cu is 56.6%, what mass of Cu is produced? *Answer* 3.70 g

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

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 of CS_2 is combined with 1.00 mol of 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 of Mg are heated with 12.0 g of 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. a. What is the limiting reactant when 19.9 g of CuO are exposed to 2.02 g of H_2 according to the following equation?

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

b. How many grams of Cu are produced?

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

CHAPTER 9 REVIEW

CHAPTER SUMMARY

9-1

- Reaction stoichiometry involves the mass relationships between reactants and products in a chemical reaction.
- 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 using mass units such as grams, kilograms, and 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.

Vocabulary

composition stoichiometry (275) mole ratio (276) reaction stoichiometry (275)

9-2

- 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 are known.

9-3

- In actual reactions, the reactants are usually combined in proportions different from the precise proportions required for complete reaction.
- The limiting reactant controls the maximum possible amount of product formed.
- Given certain quantities of reactants, the quantity of the product is always less than the maximum possible. Percent yield shows the relationship between the theoretical yield and actual yield for the product of a reaction.

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

Vocabulary

actual yield (293) limiting reactant (288) percent yield (293) theoretical yield (293)
excess reactant (288)

REVIEWING CONCEPTS

1. a. Explain the concept of *mole ratio* as used in reaction-stoichiometry problems.
b. What is the source of this value? (9-1)

2. For each of the following chemical equations, write all possible mole ratios:
a. $2Ca + O_2 \longrightarrow 2CaO$
b. $Mg + 2HF \longrightarrow MgF_2 + H_2$ (9-1)

3. a. What is molar mass?
b. What is its role in reaction stoichiometry? (9-2)

4. Distinguish between ideal and real stoichiometric calculations. (9-3)

5. Distinguish between the limiting reactant and the excess reactant in a chemical reaction. (9-3)

6. a. Distinguish between the theoretical and actual yields in stoichiometric calculations.
b. How do the values of the theoretical and actual yields generally compare? (9-3)

7. What is the percent yield of a reaction? (9-3)

8. Why are actual yields generally less than those calculated theoretically? (9-3)

PROBLEMS

General Stoichiometry

Do not assume that equations without listed coefficients are balanced.

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

10. 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 9-1.)

11. 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?

12. 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?

13. Iron is generally produced from iron ore through the following reaction in a blast furnace: $Fe_2O_3(s) + CO(g) \longrightarrow Fe(s) + CO_2(g)$.
a. If 4.00 kg of Fe_2O_3 are available to react, how many moles of CO are needed?
b. How many moles of each product are formed?

14. Methanol, CH_3OH, is an important industrial compound that is produced from the following 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 9-5.)

15. 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 of O_2?
b. How many grams of NO are required to react with this amount of O_2?

16. 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 reaction: $NaOH + CO_2 \longrightarrow Na_2CO_3 + H_2O$.
a. If the average human body discharges 925.0 g of 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?

17. 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 reacts, what mass of sodium bromide is required?
b. What mass of silver bromide is formed?

18. 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?

19. 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 of NaOH?
c. What mass of each product is formed by this reaction? (Hint: See Sample Problem 9-2.)

20. 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 9-4.)

21. 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 (in kg) could be produced from 75.0 mol of salicylic acid?

b. What mass of acetic anhydride (in 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/cm^3.

Limiting Reactant

22. Given the reactant amounts specified in each chemical equation, determine the limiting reactant in each case:

a. $HCl + NaOH \longrightarrow NaCl + H_2O$
   2.0 mol, 2.5 mol

b. $Zn + 2HCl \longrightarrow ZnCl_2 + H_2$
   2.5 mol, 6.0 mol

c. $2Fe(OH)_3 + 3H_2SO_4 \longrightarrow Fe_2(SO_4)_3 + 6H_2O$
   4.0 mol, 6.5 mol

(Hint: See Sample Problem 9-6.)

23. For each reaction specified in Problem 22, determine the amount in moles of excess reactant that remains. (Hint: See Sample Problem 9-7.)

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 of N_2H_4 were available to react with 1000. kg of $(CH_3)_2N_2H_2$ and 4500. kg of 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?

Percent Yield

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

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

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

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

d. theoretical yield = 3.45 g, percent 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 of 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 of $MgCl_2$, what is the percent 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 of 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 percent 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 percent 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 evolve when 5.00 L of gypsum are heated at 110°C to produce plaster of Paris? At 110°C, the density of $CaSO_4 \bullet 2H_2O$ is 2.32 g/mL, and the density of water vapor is 0.581 g/mL.

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 that exists 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 you would take to increase your percent yield. The reaction is as follows:

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

39. Analyzing Results In the lab, you run an experiment that appears to have a percent 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.

TECHNOLOGY & LEARNING

41. Graphing Calculator Calculating Percent Yield of a Chemical Reaction

The graphing calculator can run a program that calculates the percent yield of a chemical reaction when you enter the actual yield and the theoretical yield. Using an example in which the actual yield is 38.8 g and the theoretical yield is 53.2 g, you will calculate the percent yield. First, the program will carry out the calculation. Then it will be used to make other calculations.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data, answer these questions.

Note: all answers are written with three significant figures.

a. What is the percent yield when the actual yield is 27.3 g and the theoretical yield is 44.6 g?
b. What is the percent yield when the actual yield is 5.4 g and the theoretical yield is 9.2 g?
c. What actual yield/theoretical yield pair produced the largest percent yield?

HANDBOOK SEARCH

42. 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.
a. If 3.65×10^3 kg of iron is used in a steelmaking 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 formed?

43. 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.
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)?

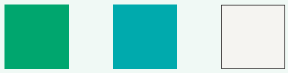
44. The reactions of oxide compounds to produce carbonates, phosphates, and sulfates are described in the section on oxides in Group 16 of the *Elements Handbook*. Use those equations to answer the following.
a. What mass of CO_2 is needed to react with 154.6 g of MgO?
b. What mass of magnesium carbonate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

RESEARCH & WRITING

45. 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

46. 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. (Hint: Don't forget to record the amount of the primary pigment and water used when you mix them.)



47. 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.