Español

CHAPTER 11 Stoichiometry

BIG Idea Mass relationships in chemical reactions confirm the law of conservation of mass.

11.1 Defining Stoichiometry

MAIN Idea The amount of each reactant present at the start of a chemical reaction determines how much product can form.

11.2 Stoichiometric Calculations

MAIN Idea The solution to every stoichiometric problem requires a balanced chemical equation.

11.3 Limiting Reactants

MAIN Idea A chemical reaction stops when one of the reactants is used up.

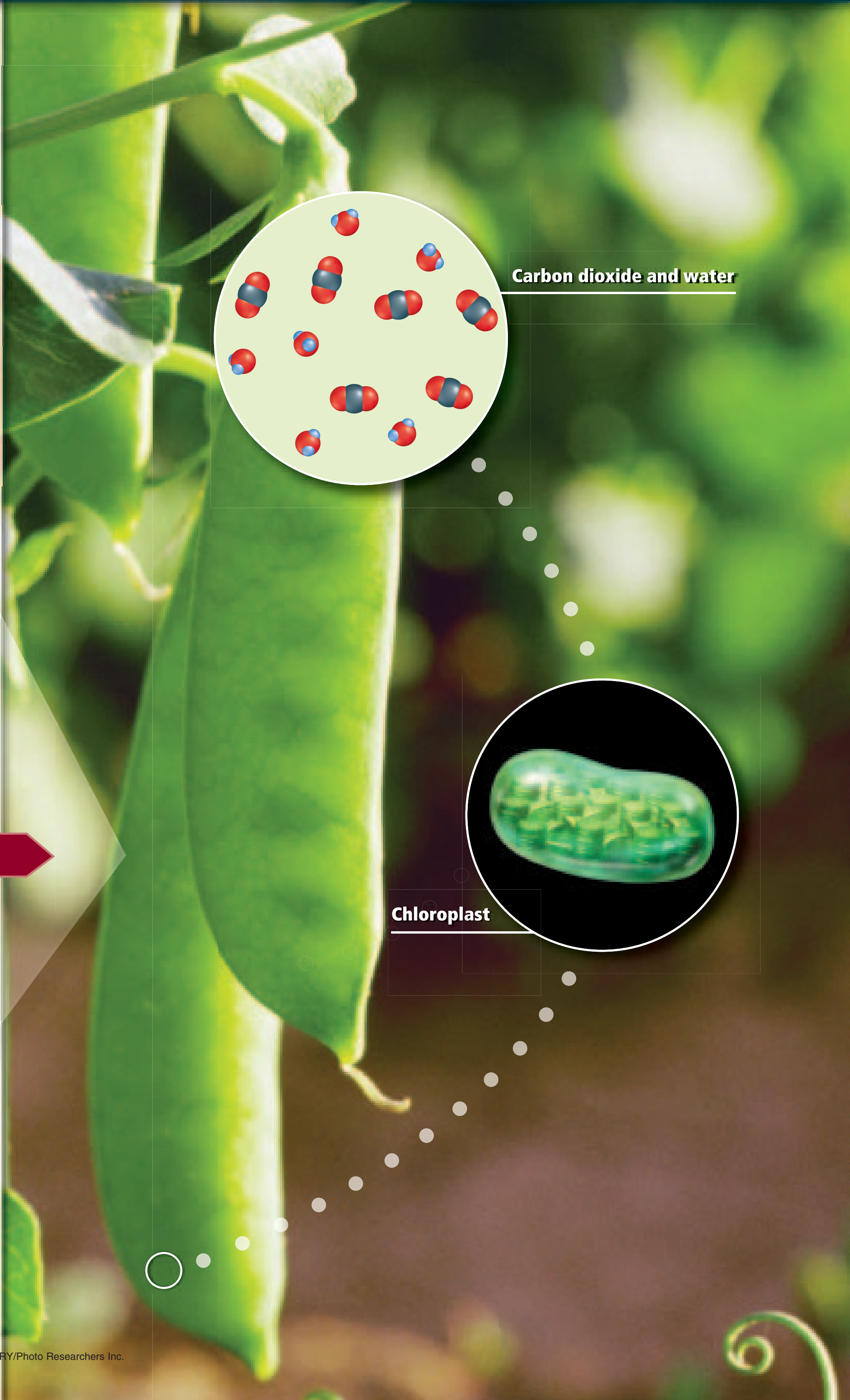
11.4 Percent Yield

MAIN Idea Percent yield is a measure of the efficiency of a chemical reaction.

ChemFacts

- Green plants make their own food through photosynthesis.
- Photosynthesis occurs within structures called chloroplasts in the cells of plants.
- The balanced chemical equation for the photosynthesis is:

  $6CO_2 + 6H_2O \longrightarrow C_6H_{12}O_6 + 6O_2$
- On a summer day, one acre of corn produces enough oxygen (a product of photosynthesis) to meet the respiratory needs of 130 people.



LAUNCH Lab

What evidence can you observe that a reaction is taking place?

During a chemical reaction, reactants are consumed as new products are formed. Often, there are several telltale signs that a chemical reaction is taking place.

Procedure

1. Read and complete the lab safety form.
2. Use a **10-mL graduated cylinder** to measure out 5.0 mL **0.01*M* potassium permanganate** ($KMnO_4$). Add the solution to a **100-mL beaker.**
3. Clean and dry the graduated cylinder, and then use it to measure 5.0 mL **0.01*M* sodium hydrogen sulfite solution** ($NaHSO_3$). Slowly add this solution to the beaker while stirring with a **stirring rod.** Record your observations.
4. Repeat Step 3 until the $KMnO_4$ solution in the beaker turns colorless. Stop adding the $NaHSO_3$ solution as soon as you obtain a colorless solution. Record your observations.

Analysis

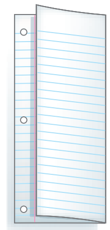
1. **Identify** the evidence you observed that a chemical reaction was occurring.
2. **Explain** why slowly adding the $NaHSO_3$ solution while stirring is a better experimental technique than adding 5.0 mL of the solution all at once.

Inquiry Would anything more have happened if you continued to add $NaHSO_3$ solution to the beaker? Explain.

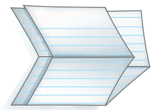
FOLDABLES™ Study Organizer

Steps in Stoichiometric Calculations Make the following Foldable to help you summarize the steps in solving a stoichiometric problem.

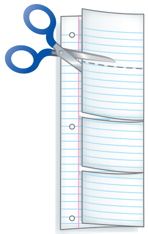
STEP 1 Fold a sheet of paper in half lengthwise.



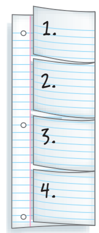
STEP 2 Fold in half widthwise and then in half again.



STEP 3 Unfold and cut along the folds of the top flap to make four tabs.



STEP 4 Label the tabs with the steps in stoichiometric calculations.



FOLDABLES Use this Foldable with Section 11.2. As you read this section, summarize each step on a tab and include an example of the step.

Chemistry Online

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Section 11.1

Objectives

- **Describe** the types of relationships indicated by a balanced chemical equation.
- **State** the mole ratios from a balanced chemical equation.

Review Vocabulary

reactant: the starting substance in a chemical reaction

New Vocabulary

stoichiometry
mole ratio

Defining Stoichiometry

MAIN Idea The amount of each reactant present at the start of a chemical reaction determines how much product can form.

Real-World Reading Link Have you ever watched a candle burning? You might have watched the candle burn out as the last of the wax was used up. Or, maybe you used a candle snuffer to put out the flame. Either way, when the candle stopped burning, the combustion reaction ended.

Particle and Mole Relationships

In doing the Launch Lab, were you surprised when the purple color of potassium permanganate disappeared as you added sodium hydrogen sulfite? If you concluded that the potassium permanganate had been used up and the reaction had stopped, you are right. Chemical reactions stop when one of the reactants is used up. When planning the reaction of potassium permanganate and sodium hydrogen sulfite, a chemist might ask, "How many grams of potassium permanganate are needed to react completely with a known mass of sodium hydrogen sulfite?" Or, when analyzing a photosynthesis reaction, you might ask, "How much oxygen and carbon dioxide are needed to form a known mass of sugar."
Stoichiometry is the tool for answering these questions.

Stoichiometry The study of quantitative relationships between the amounts of reactants used and amounts of products formed by a chemical reaction is called **stoichiometry.** Stoichiometry is based on the law of conservation of mass. Recall from Chapter 3 that the law states that matter is neither created nor destroyed in a chemical reaction. In any chemical reaction, the amount of matter present at the end of the reaction is the same as the amount of matter present at the beginning. Therefore, the mass of the reactants equals the mass of the products. Note the reaction of powdered iron (Fe) with oxygen (O_2) shown in **Figure 11.1.** Although iron reacts with oxygen to form a new compound, iron(III) oxide (Fe_2O_3), the total mass is unchanged.

■ **Figure 11.1** The balanced chemical equation for this reaction between iron and oxygen provides the relationships between amounts of reactants and products.

Concepts In Motion

Interactive Table Explore balanced chemical equations at glencoe.com.

Table 11.1 Relationships Derived from a Balanced Chemical Equation

$4Fe(s)$	+	$3O_2(g)$	$\longrightarrow$	$2Fe_2O_3(s)$
iron	+	oxygen	$\longrightarrow$	iron(III) oxide
4 atoms Fe	+	3 molecules O_2	$\longrightarrow$	2 formula units Fe_2O_3
4 mol Fe	+	3 mol O_2	$\longrightarrow$	2 mol Fe_2O_3
223.4 g Fe	+	96.00 g O_2	$\longrightarrow$	319.4 g Fe_2O_3
319.4 g reactants			$\longrightarrow$	319.4 g products

The balanced chemical equation for the chemical reaction shown in **Figure 11.1** is as follows.

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

You can interpret this equation in terms of representative particles by saying that four atoms of iron react with three molecules of oxygen to produce two formula units of iron(III) oxide. Remember that coefficients in an equation represent not only numbers of individual particles but also numbers of moles of particles. Therefore, you can also say that four moles of iron react with three moles of oxygen to produce two moles of iron(III) oxide.

The chemical equation does not directly tell you anything about the masses of the reactants and products. However, by converting the known mole quantities to mass, the mass relationships become obvious. Recall that moles are converted to mass by multiplying by the molar mass. The masses of the reactants are as follows.

$$4\ \cancel{\text{mol Fe}} \times \frac{55.85\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = 223.4\ \text{g Fe}$$

$$3\ \cancel{\text{mol O}_2} \times \frac{32.00\ \text{g O}_2}{1\ \cancel{\text{mol O}_2}} = 96.00\ \text{g O}_2$$

The total mass of the reactants is: (223.4 g + 96.00 g) = 319.4 g

Similarly, the mass of the product is calculated as follows:

$$2\ \cancel{\text{mol Fe}_2\text{O}_3} \times \frac{159.7\ \text{g Fe}_2\text{O}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} = 319.4\ \text{g}$$

Note that the mass of the reactants equals the mass of the product.

$$\text{mass of reactants} = \text{mass of products}$$
$$319.4\ \text{g} = 319.4\ \text{g}$$

As predicted by the law of conservation of mass, the total mass of the reactants equals the mass of the product. The relationships that can be determined from a balanced chemical equation are summarized in **Table 11.1.**

Reading Check **List** the types of relationships that can be derived from the coefficients in a balanced chemical equation.

VOCABULARY

WORD ORIGIN

Stoichiometry
comes from the Greek words *stoikheion,* which means element, and *metron,* which means to measure

EXAMPLE Problem 11.1

Math Handbook
Rounding
page 952

Interpreting Chemical Equations The combustion of propane (C_3H_8) provides energy for heating homes, cooking food, and soldering metal parts. Interpret the equation for the combustion of propane in terms of representative particles, moles, and mass. Show that the law of conservation of mass is observed.

1 Analyze the Problem

The coefficients in the balanced chemical equation shown below represent both moles and representative particles, in this case molecules. Therefore, the equation can be interpreted in terms of molecules and moles. The law of conservation of mass will be verified if the masses of the reactants and products are equal.

Known

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

Unknown

Equation interpreted in terms of molecules = ?

Equation interpreted in terms of moles = ?

Equation interpreted in terms of mass = ?

2 Solve for the Unknown

The coefficients in the chemical equation indicate the number of molecules.

1 molecule C_3H_8 + 5 molecules O_2 $\longrightarrow$ 3 molecules CO_2 + 4 molecules H_2O

The coefficients in the chemical equation also indicate the number of moles.

1 mol C_3H_8 + 5 mol O_2 $\longrightarrow$ 3 mol CO_2 + 4 mol H_2O

To verify that mass is conserved, first convert moles of reactant and product to mass by multiplying by a conversion factor—the molar mass—that relates grams to moles.

$$\text{moles of reactant or product} \times \frac{\text{grams reactant or product}}{\text{1 mol reactant or product}} = \text{grams of reactant or product}$$

$$1\ \cancel{\text{mol}\ C_3H_8} \times \frac{44.09\ \text{g}\ C_3H_8}{1\ \cancel{\text{mol}\ C_3H_8}} = 44.09\ \text{g}\ C_3H_8$$

Calculate the mass of the reactant C_3H_8.

$$5\ \cancel{\text{mol}\ O_2} \times \frac{32.00\ \text{g}\ O_2}{1\ \cancel{\text{mol}\ O_2}} = 160.0\ \text{g}\ O_2$$

Calculate the mass of the reactant O_2.

$$3\ \cancel{\text{mol}\ CO_2} \times \frac{44.01\ \text{g}\ CO_2}{1\ \cancel{\text{mol}\ CO_2}} = 132.0\ \text{g}\ CO_2$$

Calculate the mass of the product CO_2.

$$4\ \cancel{\text{mol}\ H_2O} \times \frac{18.02\ \text{g}\ H_2O}{1\ \cancel{\text{mol}\ H_2O}} = 72.08\ \text{g}\ H_2O$$

Calculate the mass of the product H_2O

44.09 g C_3H_8 + 160.0 g O_2 = **204.1 g reactants** — **Add the masses of the reactants.**

132.0 g CO_2 + 72.08 g H_2O = **204.1 g products** — **Add the masses of the products.**

204.1 g reactants = 204.1 g products — **The law of conservation of mass is observed.**

3 Evaluate the Answer

The sums of the reactants and the products are correctly stated to the first decimal place because each mass is accurate to the first decimal place. The mass of reactants equals the mass of products, as predicted by the law of conservation of mass.

PRACTICE Problems

Extra Practice Pages 982–983 and glencoe.com

1. Interpret the following balanced chemical equations in terms of particles, moles, and mass. Show that the law of conservation of mass is observed.
   a. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$
   b. $HCl(aq) + KOH(aq) \longrightarrow KCl(aq) + H_2O(l)$
   c. $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
2. **Challenge** For each of the following, balance the chemical equation; interpret the equation in terms of particles, moles, and mass; and show that the law of conservation of mass is observed.
   a. $\_\_Na(s) + \_\_H_2O(l) \longrightarrow \_\_NaOH(aq) + \_\_H_2(g)$
   b. $\_\_Zn(s) + \_\_HNO_3(aq) \longrightarrow \_\_Zn(NO_3)_2(aq) + \_\_N_2O(g) + \_\_H_2O(l)$

Mole ratios You have read that the coefficients in a chemical equation indicate the relationships between moles of reactants and products. You can use the relationships between coefficients to derive conversion factors called mole ratios. A **mole ratio** is a ratio between the numbers of moles of any two of the substances in a balanced chemical equation. For example, consider the reaction shown in **Figure 11.2.** In this reaction, potassium (K) reacts with bromine (Br_2) to form potassium bromide (KBr). The product of the reaction, the ionic salt potassium bromide, is prescribed by veterinarians as an antiepileptic medication for dogs and cats.

$$2K(s) + Br_2(l) \longrightarrow 2KBr(s)$$

What mole ratios can be written for this reaction? Starting with the reactant potassium, you can write a mole ratio that relates the moles of potassium to each of the other two substances in the equation. Thus, one mole ratio relates the moles of potassium used to the moles of bromine used. The other mole ratio relates the moles of potassium used to the moles of potassium bromide formed.

$$\frac{2\text{ mol K}}{1\text{ mol Br}_2} \text{ and } \frac{2\text{ mol K}}{2\text{ mol KBr}}$$

Two other mole ratios show how the moles of bromine relate to the moles of the other two substances in the equation—potassium and potassium bromide.

$$\frac{1\text{ mol Br}_2}{2\text{ mol K}} \text{ and } \frac{1\text{ mol Br}_2}{2\text{ mol KBr}}$$

Similarly, two ratios relate the moles of potassium bromide to the moles of potassium and bromine.

$$\frac{2\text{ mol KBr}}{2\text{ mol K}} \text{ and } \frac{2\text{ mol KBr}}{1\text{ mol Br}_2}$$

These six ratios define all the mole relationships in this equation. Each of the three substances in the equation forms a ratio with the two other substances.

Reading Check **Identify** the source from which a chemical reaction's mole ratios are derived.

Chemistry Online
Personal Tutor For an online tutorial on ratios, visit glencoe.com.

Figure 11.2 Potassium metal and liquid bromine react vigorously to form the ionic compound potassium bromide. Bromine is one of the two elements that are liquids at room temperature (mercury is the other). Potassium is a highly reactive metal.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

3. Determine all possible mole ratios for the following balanced chemical equations.

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

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

c. $2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$

4. Challenge Balance the following equations, and determine the possible mole ratios.

a. $ZnO(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2O(l)$

b. butane (C_4H_{10}) + oxygen $\longrightarrow$ carbon dioxide + water

The decomposition of potassium chlorate ($KClO_3$) is sometimes used to obtain small amounts of oxygen in the laboratory.

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

The mole ratios that can be written for this reaction are as follows.

$$\frac{2 \text{ mol } KClO_3}{2 \text{ mol } KCl} \text{ and } \frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$$

$$\frac{2 \text{ mol } KCl}{2 \text{ mol } KClO_3} \text{ and } \frac{2 \text{ mol } KCl}{3 \text{ mol } O_2}$$

$$\frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} \text{ and } \frac{3 \text{ mol } O_2}{2 \text{ mol } KCl}$$

Note that the number of mole ratios you can write for a chemical reaction involving a total of n substances is $(n)(n-1)$. Thus, for reactions involving four and five substances, you can write 12 and 20 moles ratios, respectively.

$$\text{Four substances: } (4)(3) = 12 \text{ mole ratios}$$
$$\text{Five substances: } (5)(4) = 20 \text{ mole ratios}$$

VOCABULARY

ACADEMIC VOCABULARY

Derive
to obtain from a specified source
The researcher was able to derive the meaning of the illustration from ancient texts.

Section 11.1 Assessment

Section Summary

- Balanced chemical equations can be interpreted in terms of moles, mass, and representative particles (atoms, molecules, formula units).
- The law of conservation of mass applies to all chemical reactions.
- Mole ratios are derived from the coefficients of a balanced chemical equation. Each mole ratio relates the number of moles of one reactant or product to the number of moles of another reactant or product in the chemical reaction.

5. MAIN Idea **Compare** the mass of the reactants and the mass of the products in a chemical reaction, and explain how these masses are related.

6. State how many mole ratios can be written for a chemical reaction involving three substances.

7. Categorize the ways in which a balanced chemical equation can be interpreted.

8. Apply The general form of a chemical reaction is $xA + yB \longrightarrow zAB$. In the equation, A and B are elements, and x, y, and z are coefficients. State the mole ratios for this reaction.

9. Apply Hydrogen peroxide (H_2O_2) decomposes to produce water and oxygen. Write a balanced chemical equation for this reaction, and determine the possible mole ratios.

10. Model Write the mole ratios for the reaction of hydrogen gas and oxygen gas, $2H_2(g) + O_2(g) \longrightarrow 2H_2O$. Make a sketch of six hydrogen molecules reacting with the correct number of oxygen molecules. Show the water molecules produced.

Chemistry Online Self-Check Quiz glencoe.com

Section 11.2

Objectives

- **List** the sequence of steps used in solving stoichiometric problems.
- **Solve** stoichiometric problems.

Review Vocabulary

chemical reaction: a process in which the atoms of one or more substances are rearranged to form different substances

Stoichiometric Calculations

MAIN Idea The solution to every stoichiometric problem requires a balanced chemical equation.

Real-World Reading Link Baking requires accurate measurements. That is why it is necessary to follow a recipe when baking cookies from scratch. If you need to make more cookies than a recipe yields, what must you do?

Using Stoichiometry

What tools are needed to perform stoichiometric calculations? All stoichiometric calculations begin with a balanced chemical equation. Mole ratios based on the balanced chemical equation are needed, as well as mass-to-mole conversions.

Stoichiometric mole-to-mole conversion The vigorous reaction between potassium and water is shown in **Figure 11.3.** The balanced chemical equation is as follows.

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$

FOLDABLES Incorporate information from this section into your Foldable.

From the balanced equation, you know that two moles of potassium yields one mole of hydrogen. But how much hydrogen is produced if only 0.0400 mol of potassium is used? To answer this question, identify the given, or known, substance and the substance that you need to determine. The given substance is 0.0400 mol of potassium. The unknown is the number of moles of hydrogen. Because the given substance is in moles and the unknown substance to be determined is also in moles, this problem involves a mole-to-mole conversion.

To solve the problem, you need to know how the unknown moles of hydrogen are related to the known moles of potassium. In Section 11.1, you learned to derive mole ratios from the balanced chemical equation. Mole ratios are used as conversion factors to convert the known number of moles of one substance to the unknown number of moles of another substance in the same reaction. Several mole ratios can be written from the equation, but how do you choose the correct one?

■ **Figure 11.3** Potassium metal reacts vigorously with water, releasing so much heat that the hydrogen gas formed in the reaction catches fire.

As shown below, the correct mole ratio, 1 mol H_2 to 2 mol K, has moles of unknown in the numerator and moles of known in the denominator. Using this mole ratio converts the moles of potassium to the unknown number of moles of hydrogen.

$$\text{moles of known} \times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$$

$$0.0400\ \cancel{\text{mol K}} \times \frac{1\ \text{mol H}_2}{2\ \cancel{\text{mol K}}} = 0.0200\ \text{mol H}_2$$

The following Example Problems show mole-to-mole, mole-to-mass, and mass-to-mass stoichiometry problems. The process used to solve these problems is outlined in the Problem-Solving Strategy below.

Problem-Solving Strategy

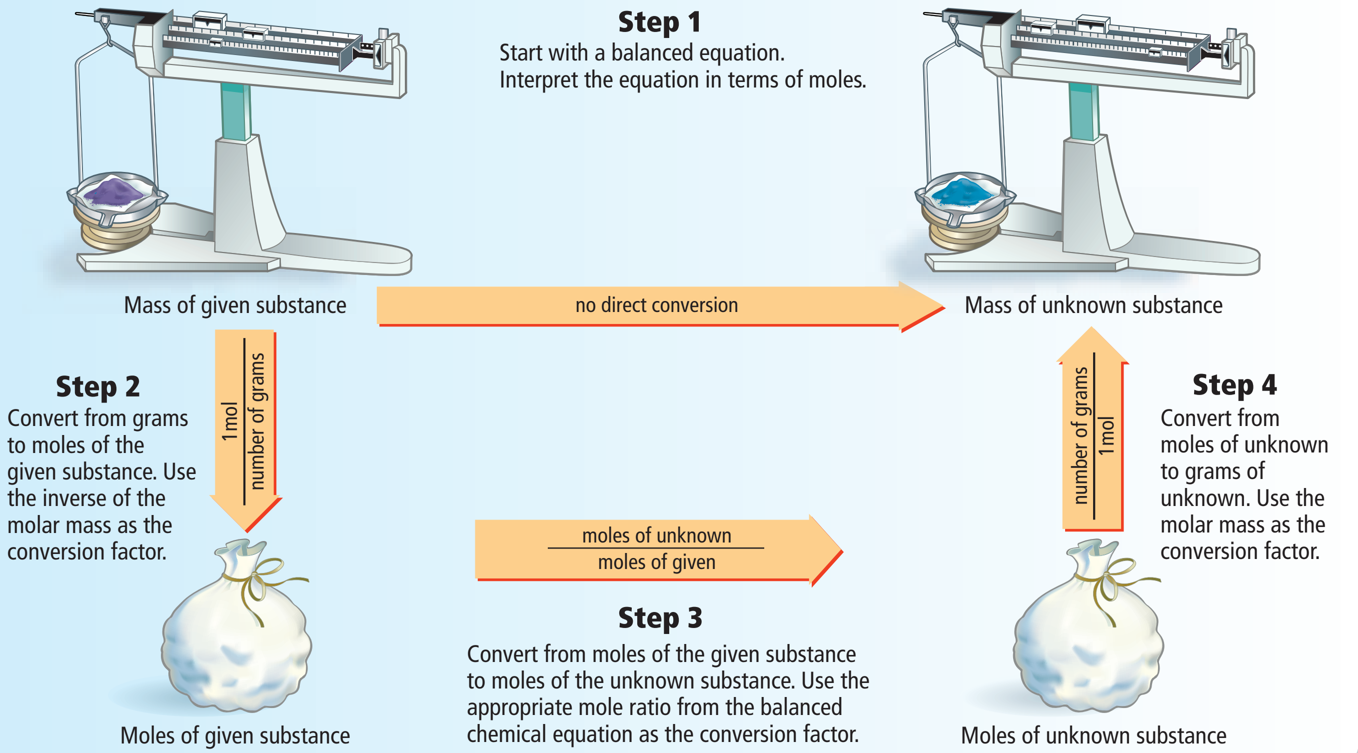
Mastering Stoichiometry

The flowchart below outlines the steps used to solve mole-to-mole, mole-to-mass, and mass-to-mass stoichiometric problems.

1. Complete Step 1 by writing the balanced chemical equation for the reaction.
2. To determine where to start your calculations, note the unit of the given substance.
   - If mass (in grams) of the given substance is the starting unit, begin your calculations with Step 2.
   - If amount (in moles) of the given substance is the starting unit, skip Step 2 and begin your calculations with Step 3.
3. The end point of the calculation depends on the desired unit of the unknown substance.
   - If the answer must be in moles, stop after completing Step 3.
   - If the answer must be in grams, stop after completing Step 4.

Apply the Strategy

Apply the Problem-Solving Strategy to Example Problems 11.2, 11.3, and 11.4.



EXAMPLE Problem 11.2

Mole-to-Mole Stoichiometry One disadvantage of burning propane (C_3H_8) is that carbon dioxide (CO_2) is one of the products. The released carbon dioxide increases the concentration of CO_2 in the atmosphere. How many moles of CO_2 is produced when 10.0 mol of C_3H_8 is burned in excess oxygen in a gas grill?

1 Analyze the Problem

You are given moles of the reactant, C_3H_8 and must find the moles of the product, CO_2. First write the balanced chemical equation, then convert from moles of C_3H_8 to moles of CO_2. The correct mole ratio has moles of unknown substance in the numerator and moles of known substance in the denominator.

Known

moles C_3H_8 = 10.0 mol C_3H_8

Unknown

moles CO_2 = ? mol CO_2

Math Handbook
Ratios
page 964

2 Solve for the Unknown

Write the balanced chemical equation for the combustion of C_3H_8. Use the correct mole ratio to convert moles of known (C_3H_8) to moles of unknown (CO_2).

$$\overset{10.0\text{ mol}}{C_3H_8(g)} + 5O_2(g) \rightarrow \overset{?\text{ mol}}{3CO_2(g)} + 4H_2O(g)$$

$$\text{Mole ratio: } \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$$

$$10.0 \cancel{\text{ mol } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \cancel{\text{ mol } C_3H_8}} = \mathbf{30.0 \text{ mol } CO_2}$$

Burning 10.0 moles of C_3H_8 produces 30.0 moles CO_2.

3 Evaluate the Answer

Because the given number of moles has three significant figures, the answer also has three figures. The balanced chemical equation indicates that 1 mol of C_3H_8 produces 3 mol of CO_2. Thus, 10.0 mol of C_3H_8 produces three times as many moles of CO_2, or 30.0 mol.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

11. Methane and sulfur react to produce carbon disulfide (CS_2), a liquid often used in the production of cellophane.

$$\_\_CH_4(g) + \_\_S_8(s) \rightarrow \_\_CS_2(l) + \_\_H_2S(g)$$

a. Balance the equation.

b. Calculate the moles of CS_2 produced when 1.50 mol S_8 is used.

c. How many moles of H_2S is produced?

12. Challenge Sulfuric acid (H_2SO_4) is formed when sulfur dioxide (SO_2) reacts with oxygen and water.

a. Write the balanced chemical equation for the reaction.

b. How many moles of H_2SO_4 is produced from 12.5 moles of SO_2?

c. How many moles of O_2 are needed?

Real-World Chemistry
Outdoor Cooking

Gas Grills Using outdoor grills is a popular way to cook. Gas grills burn either natural gas or propane that is mixed with air. The initial spark is provided by a grill starter. Propane is more commonly used for fuel because it can be supplied in liquid form in a portable tank. Combustion of liquid propane also releases more energy than natural gas.

Stoichiometric mole-to-mass conversion Now, suppose you know the number of moles of a reactant or product in a reaction and you want to calculate the mass of another product or reactant. This is an example of a mole-to-mass conversion.

EXAMPLE Problem 11.3

Math Handbook
Calculations with Significant Figures
pages 952–953

Mole-to-Mass Stoichiometry Determine the mass of sodium chloride (NaCl), commonly called table salt, produced when 1.25 mol of chlorine gas (Cl_2) reacts vigorously with excess sodium.

1 Analyze the Problem

You are given the moles of the reactant, Cl_2, and must determine the mass of the product, NaCl. You must convert from moles of Cl_2 to moles of NaCl using the mole ratio from the equation. Then, you need to convert moles of NaCl to grams of NaCl using the molar mass as the conversion factor.

Known
moles of chlorine = 1.25 mol Cl_2

Unknown
mass of sodium chloride = ? g NaCl

2 Solve for the Unknown

$$\underset{}{2Na(s)} + \overset{1.25\text{ mol}}{Cl_2(g)} \longrightarrow \overset{?\text{ g}}{2NaCl(s)}$$

Write the balanced chemical equation, and identify the known and the unknown values.

Mole ratio: $\dfrac{2 \text{ mol NaCl}}{1 \text{ mol } Cl_2}$

$$1.25 \cancel{\text{mol } Cl_2} \times \frac{2 \text{ mol NaCl}}{1 \cancel{\text{mol } Cl_2}} = 2.50 \text{ mol NaCl}$$

Multiply moles of Cl_2 by the mole ratio to get moles of NaCl.

$$2.50 \cancel{\text{mol NaCl}} \times \frac{58.44 \text{ g NaCl}}{1 \cancel{\text{mol NaCl}}} = \mathbf{146 \text{ g NaCl}}$$

Multiply moles of NaCl by the molar mass to get grams of NaCl.

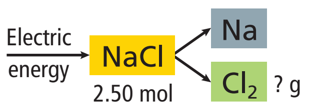
3 Evaluate the Answer

Because the given number of moles has three significant figures, the mass of NaCl also has three. To quickly assess whether the calculated mass value for NaCl is correct, perform the calculations in reverse: divide the mass of NaCl by the molar mass of NaCl, and then divide the result by 2. You will obtain the given number of moles of Cl_2.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

13. Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How much chlorine gas, in grams, is obtained from the process diagrammed at right?



14. Challenge Titanium is a transition metal used in many alloys because it is extremely strong and lightweight. Titanium tetrachloride ($TiCl_4$) is extracted from titanium oxide (TiO_2) using chlorine and coke (carbon).

$$TiO_2(s) + C(s) + 2Cl_2(g) \longrightarrow TiCl_4(s) + CO_2(g)$$

a. What mass of Cl_2 gas is needed to react with 1.25 mol of TiO_2?

b. What mass of C is needed to react with 1.25 mol of TiO_2?

c. What is the mass of all of the products formed by reaction with 1.25 mol of TiO_2?

Stoichiometric mass-to-mass conversion If you were preparing to carry out a chemical reaction in the laboratory, you would need to know how much of each reactant to use in order to produce the mass of product you required. Example Problem 11.4 demonstrates how you can use a measured mass of the known substance, the balanced chemical equation, and mole ratios from the equation to find the mass of the unknown substance. The ChemLab at the end of this chapter will provide you with laboratory experience in determining a mole ratio.

EXAMPLE Problem 11.4

Math Handbook
Dimensional Analysis
page 956

Mass-to-Mass Stoichiometry Ammonium nitrate (NH_4NO_3), an important fertilizer, produces dinitrogen oxide (N_2O) gas and H_2O when it decomposes. Determine the mass of H_2O produced from the decomposition of 25.0 g of solid NH_4NO_3.

1 Analyze the Problem

You are given a description of the chemical reaction and the mass of the reactant. You need to write the balanced chemical equation and convert the known mass of the reactant to moles of the reactant. Then, use a mole ratio to relate moles of the reactant to moles of the product. Finally, use the molar mass to convert from moles of the product to the mass of the product.

Known
mass of ammonium nitrate = 25.0 g NH_4NO_3

Unknown
mass of water = ? g H_2O

2 Solve for the Unknown

$$\overset{25.0\text{ g}}{NH_4NO_3(s)} \rightarrow N_2O(g) + \overset{?\text{ g}}{2H_2O(g)}$$

Write the balanced chemical equation, and identify the known and unknown values.

$$25.0\ \cancel{\text{g } NH_4NO_3} \times \frac{1\text{ mol } NH_4NO_3}{80.04\ \cancel{\text{g } NH_4NO_3}} = 0.312\text{ mol } NH_4NO_3$$

Multiply grams of NH_4NO_3 by the inverse of molar mass to get moles of NH_4NO_3.

$$\text{Mole ratio: } \frac{2\text{ mol H2O}}{1\text{ mol } NH_4NO_3}$$

$$0.312\ \cancel{\text{mol } NH_4NO_3} \times \frac{2\text{ mol } H_2}{1\ \cancel{\text{mol } NH_4NO_3}} = 0.624\text{ mol } H_2O$$

Multiply moles of NH_4NO_3 by the mole ratio to get moles of H_2O.

$$0.624\ \cancel{\text{mol } H_2O} \times \frac{18.02\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}} = \mathbf{11.2\text{ g } H_2O}$$

Multiply moles of H_2O by the molar mass to get grams of H_2O.

3 Evaluate the Answer

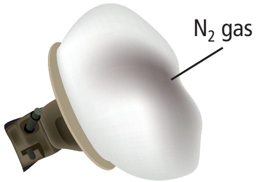
The number of significant figures in the answer, three, is determined by the given moles of NH_4NO_3. To verify that the mass of H_2O is correct, perform the calculations in reverse.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

15. One of the reactions used to inflate automobile air bags involves sodium azide (NaN_3): $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$. Determine the mass of N_2 produced from the decomposition of NaN_3 shown at right.

16. Challenge In the formation of acid rain, sulfur dioxide (SO_2) reacts with oxygen and water in the air to form sulfuric acid (H_2SO_4). Write the balanced chemical equation for the reaction. If 2.50 g of SO_2 reacts with excess oxygen and water, how much H_2SO_4, in grams, is produced?



100.0 g $NaN_3 \rightarrow$? g $N_2(g)$

MiniLab

Apply Stoichiometry

How much sodium carbonate (Na_2CO_3) is produced when baking soda decomposes? Baking soda is used in many baking recipes because it makes batter rise, which results in a light and fluffy texture. This occurs because baking soda, sodium hydrogen carbonate ($NaHCO_3$), decomposes upon heating to form carbon dioxide gas according to the following equation.

$$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$$

Procedure

1. Read and complete the lab safety form.
2. Create a data table to record your experimental data and observation.
3. Use a **balance** to measure the mass of a clean, dry **crucible.** Add about 3.0 g of **sodium hydrogen carbonate ($NaHCO_3$),** and measure the combined mass of the crucible and $NaHCO_3$. Record both masses in your data table, and calculate the mass of the $NaHCO_3$.
4. Use this starting mass of $NaHCO_3$ and the balanced chemical equation to calculate the mass of $NaHCO_3$ that will be produced.
5. Set up a **ring stand** with a **ring** and **clay triangle** for heating the crucible.
6. Heat the crucible with a **Bunsen burner,** slowly at first and then with a stronger flame, for 7–8 min. Record your observations during the heating.
7. Turn off the burner, and use **crucible tongs** to remove the hot crucible.

   WARNING: ***Do not touch the hot crucible with your hands.***
8. Allow the crucible to cool, and then measure the mass of the crucible and $NaHCO_3$.

Analysis

1. **Describe** what you observed during the heating of the baking soda.
2. **Compare** your calculated mass of $NaHCO_3$ with the actual mass you obtained from the experiment.
3. **Calculate** Assume that the mass of Na_2HCO_3 that you calculated in Step 4 is the accepted value for the mass of product that will form. Calculate the error and percent error associated with the experimentally measured mass.
4. **Identify** sources of error in the procedure that led to errors calculated in Question 3.

Section 11.2 Assessment

Section Summary

- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.
- Stoichiometric problems make use of mole ratios to convert between mass and moles.

17. MAIN Idea **Explain** why a balanced chemical equation is needed to solve a stoichiometric problem.
18. **List** the four steps used in solving stoichiometric problems.
19. **Describe** how a mole ratio is correctly expressed when it is used to solve a stoichiometric problem.
20. **Apply** How can you determine the mass of liquid bromine (Br_2) needed to react completely with a given mass of magnesium?
21. **Calculate** Hydrogen reacts with excess nitrogen as follows:

    $$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

    If 2.70 g of H_2 reacts, how many grams of NH_3 is formed?
22. **Design** a concept map for the following reaction.

    $$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

    The concept map should explain how to determine the mass of $CaCl_2$ produced from a given mass of HCl.

Chemistry Online Self-Check Quiz glencoe.com

Section 11.3

Objectives

- **Identify** the limiting reactant in a chemical equation.
- **Identify** the excess reactant, and calculate the amount remaining after the reaction is complete.
- **Calculate** the mass of a product when the amounts of more than one reactant are given.

Review Vocabulary

molar mass: the mass in grams of one mole of any pure substance

New Vocabulary

limiting reactant
excess reactant

Limiting Reactants

MAIN Idea A chemical reaction stops when one of the reactants is used up.

Real-World Reading Link If there are more boys than girls at a school dance, some boys will be left without dance partners. The situation is much the same for the reactants in a chemical reaction—excess reactants cannot participate.

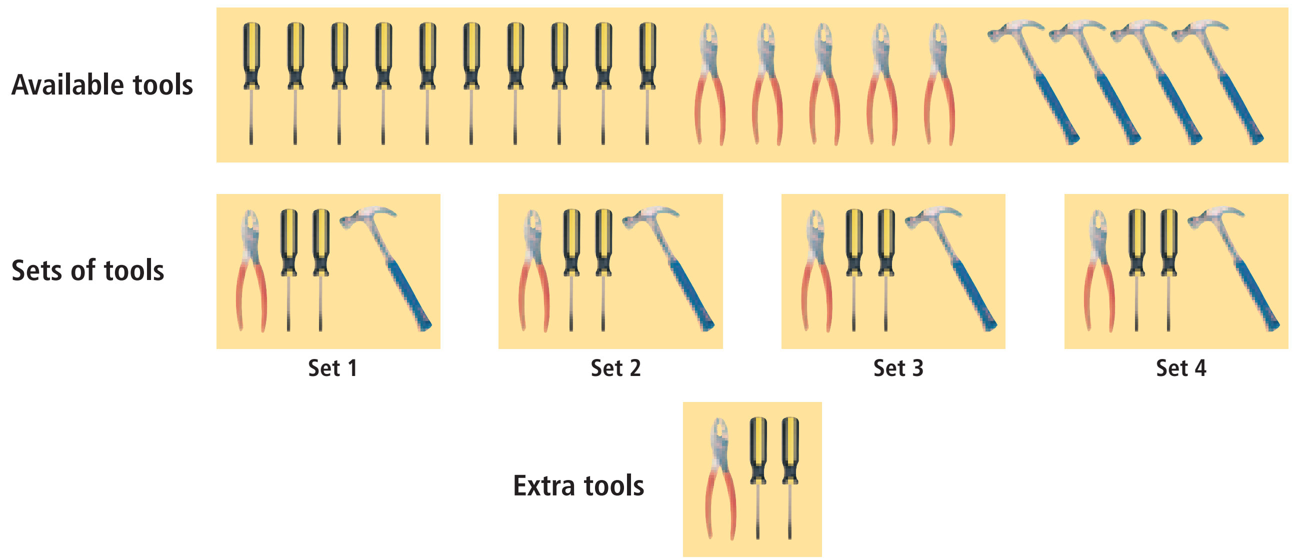
Why do reactions stop?

Rarely in nature are the reactants present in the exact ratios specified by the balanced chemical equation. Generally, one or more reactants are in excess and the reaction proceeds until all of one reactant is used up. When a reaction is carried out in the laboratory, the same principle applies. Usually, one or more of the reactants are in excess, while one is limited. The amount of product depends on the reactant that is limited.

Limiting and excess reactants Recall the reaction from the Launch Lab. After the colorless solution formed, adding more sodium hydrogen sulfite had no effect because there was no more potassium permanganate available to react with it. Potassium permanganate was a limiting reactant. As the name implies, the **limiting reactant** limits the extent of the reaction and, thereby, determines the amount of product formed. A portion of all the other reactants remains after the reaction stops. Reactants leftover when a reaction stops are **excess reactants.**

To help you understand limiting and excess reactants, consider the analogy in **Figure 11.4.** From the available tools, four complete sets consisting of a pair of pliers, a hammer, and two screwdrivers can be assembled. The number of sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess.

Figure 11.4 Each tool set must have one hammer, so only four sets can be assembled.
Interpret *How many more hammers are required to complete a fifth set?*



Aaron Haupt

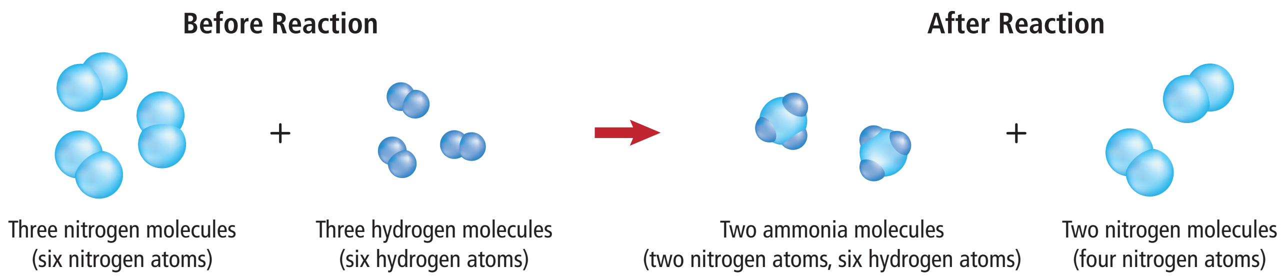


Figure 11.5 If you check all the atoms present before and after the reaction, you will find that some of the nitrogen molecules are unchanged. These nitrogen molecules are the excess reactant.

CLICK HERE

Concepts In Motion

Interactive Figure To see an animation of limiting reactants, visit glencoe.com.

Determining the limiting reactant The calculations you did in the previous section were based on having the reactants present in the ratio described by the balanced chemical equation. When this is not the case, the first thing you must do is determine which reactant is limiting.

Consider the reaction shown in **Figure 11.5,** in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). In the first step of the reaction, all the nitrogen molecules and hydrogen molecules are separated into individual atoms. These atoms are available for reassembling into ammonia molecules, just as the tools in **Figure 11.4** are available to be assembled into tool kits. How many molecules of ammonia can be produced from the available atoms? Two ammonia molecules can be assembled from the hydrogen atoms and nitrogen atoms because only six hydrogen atoms are available—three for each ammonia molecule. When the hydrogen is gone, two unreacted molecules of nitrogen remain. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It is important to know which reactant is the limiting reactant because, as you have just read, the amount of product formed depends on this reactant.

Reading Check **Extend** How many more hydrogen molecules would be needed to completely react with the excess nitrogen molecules shown in **Figure 11.5?**

Figure 11.6 Natural rubber, which is soft and very sticky, is hardened in a chemical process called vulcanization. During vulcanization, molecules become linked together, forming a durable material that is harder, smoother, and less sticky. These properties make vulcanized rubber ideal for many products, such as this caster.

Calculating the Product when a Reactant Is Limiting

How can you calculate the amount of product formed when one of the reactants is limiting? Consider the formation of disulfur dichloride (S_2Cl_2), which is used to vulcanize rubber. As shown in **Figure 11.6,** the properties of vulcanized rubber make it useful for many products. In the production of disulfur dichloride, molten sulfur reacts with chlorine gas according to the following equation.

$$S_8(l) + 4Cl_2(g) \longrightarrow 4S_2Cl_2(l)$$

If 200.0 g of sulfur reacts with 100.0 g of chlorine, what mass of disulfur dichloride is produced?

Calculating the limiting reactant The masses of both reactants are given. First, determine which one is the limiting reactant, because the reaction stops producing product when the limiting reactant is used up.

Moles of reactants Identifying the limiting reactant involves finding the number of moles of each reactant. You can do this by converting the masses of chlorine and sulfur to moles. Multiply each mass by a conversion factor that relates moles and mass—the inverse of molar mass.

$$100.0\ \cancel{g\ Cl_2} \times \frac{1\ mol\ Cl_2}{70.91\ \cancel{g\ Cl_2}} = 1.410\ mol\ Cl_2$$

$$200.0\ \cancel{g\ S_8} \times \frac{1\ mol\ S_8}{256.5\ \cancel{g\ S_8}} = 0.7797\ mol\ S_8$$

Using mole ratios The next step involves determining whether the two reactants are in the correct mole ratio, as given in the balanced chemical equation. The coefficients in the balanced chemical equation indicate that 4 mol of chlorine is needed to react with 1 mol of sulfur. This 4:1 ratio from the equation must be compared with the actual ratio of the moles of available reactants just calculated above. To determine the actual ratio of moles, divide the number of available moles of chlorine by the number of available moles of sulfur.

$$\frac{1.410\ mol\ Cl_2\ available}{0.7797\ mol\ S_8\ available} = \frac{1.808\ mol\ Cl_2\ available}{1\ mol\ S_8\ available}$$

Only 1.808 mol of chlorine is available for every 1 mol of sulfur, instead of the 4 mol of chlorine required by the balanced chemical equation. Therefore, chlorine is the limiting reactant.

Calculating the amount of product formed After determining the limiting reactant, the amount of product in moles can be calculated by multiplying the given number of moles of the limiting reactant (1.410 mol Cl_2) by the mole ratio relating disulfur dichloride and chlorine. Then, moles of S_2Cl_2 is converted to grams of S_2Cl_2 by multiplying by the molar mass. These calculations can be combined as shown.

$$1.410\ \cancel{mol\ Cl_2} \times \frac{4\ \cancel{mol\ S_2Cl_2}}{4\ \cancel{mol\ Cl_2}} \times \frac{135.0\ g\ S_2Cl_2}{1\ \cancel{mol\ S_2Cl_2}} = 190.4\ g\ S_2Cl_2$$

Thus, 190.4 g S_2Cl_2 forms when 1.410 mol Cl_2 reacts with excess S_8.

Analyzing the excess reactant Now that you have determined the limiting reactant and the amount of product formed, what about the excess reactant, sulfur? How much of it reacted?

Moles reacted You need to make a mole-to-mass calculation to determine the mass of sulfur needed to react completely with 1.410 mol of chlorine. First, obtain the number of moles of sulfur by multiplying the moles of chlorine by the S_8-to-Cl_2 mole ratio.

$$1.410\ \cancel{mol\ Cl_2} \times \frac{1\ mol\ S_8}{4\ \cancel{mol\ Cl_2}} = 0.3525\ mol\ S_8$$

Mass reacted Next, to obtain the mass of sulfur needed, multiply 0.3525 mol of S_8 by its molar mass.

$$0.3525\ \cancel{mol\ S_8} \times \frac{265.5\ g\ S_8}{1\ \cancel{mol\ S_8}} = 90.42\ g\ S_8\ needed$$

Excess remaining Knowing that 200.0 g of sulfur is available and that only 90.42 g of sulfur is needed, you can calculate the amount of sulfur left unreacted when the reaction ends.

$$200.0\ g\ S_8\ available - 90.42\ g\ S_8\ needed = 109.6\ g\ S_8\ in\ excess$$

CAREERS IN CHEMISTRY

Pharmacist Knowledge of drug composition, modes of action, and possible harmful interactions with other substances allows a pharmacist to counsel patients on their care. Pharmacists also mix chemicals to form powders, tablets, ointments, and solutions. For more information on chemistry careers, visit glencoe.com.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Product

Science usage: a new substance formed during a chemical reaction
The sole reaction product was a colorless gas.

Common usage: something produced
The cosmetics counter in the department store had hundreds of products from which to choose.

EXAMPLE Problem 11.5

Math Handbook
Dimensional Analysis
page 956

Determining the Limiting Reactant The reaction between solid white phosphorus (P_4) and oxygen produces solid tetraphosphorus decoxide (P_4O_{10}). This compound is often called diphosphorus pentoxide because its empirical formula is P_2O_5.

a. Determine the mass of P_4O_{10} formed if 25.0 g of P_4 and 50.0 g of oxygen are combined.

b. How much of the excess reactant remains after the reaction stops?

1 Analyze the Problem

You are given the masses of both reactants, so you must identify the limiting reactant and use it to find the mass of the product. From moles of the limiting reactant, the moles of the excess reactant used in the reaction can be determined. The number of moles of the excess reactant that reacted can be converted to mass and subtracted from the given mass to find the amount in excess.

Known	**Unknown**
mass of phosphorus = 25.0 g P_4	**mass of tetraphosphorus decoxide = ? g P_4O_{10}**
mass of oxygen = 50.0 g O_2	**mass of excess reactant = ? g excess reactant**

2 Solve for the Unknown

Determine the limiting reactant.

$$\underset{25.0\text{ g}}{P_4(s)} + \underset{50.0\text{ g}}{5O_2(g)} \rightarrow \underset{?\text{ g}}{P_4O_{10}(s)}$$

Write the balanced chemical equation, and identify the known and the unknown.

Determine the number of moles of the reactants by multiplying each mass by the conversion factor that relates moles and mass—the inverse of molar mass.

$$25.0\ \cancel{g\ P_4} \times \frac{1\text{ mol }P_4}{123.9\ \cancel{g\ P_4}} = 0.202\text{ mol }P_4$$

Calculate the moles of P_4.

$$50.0\ \cancel{g\ O_2} \times \frac{1\text{ mol }O_2}{32.00\ \cancel{g\ O_2}} = 1.56\text{ mol }O_2$$

Calculate the moles of O_2.

Calculate the actual ratio of available moles of O_2 and available moles of P_4.

$$\frac{1.56\text{ mol }O_2}{0.202\text{ mol }P_4} = \frac{7.72\text{ mol }O_2}{1\text{ mol }P_4}$$

Calculate the ratio of moles of O_2 to moles of P_4.

Determine the mole ratio of the two reactants from the balanced chemical equation.

Mole ratio: $\frac{5\text{ mol }O_2}{\text{mol }P_4}$

Because 7.72 mol of O_2 is available but only 5 mol is needed to react with 1 mol of P_4, O_2 is in excess and P_4 is the limiting reactant. Use the moles of P_4 to determine the moles of P_4O_{10} that will be produced. Multiply the number of moles of P_4 by the mole ratio of P_4O_{10} (the unknown) to P_4 (the known).

$$0.202\ \cancel{mol\ P_4} \times \frac{1\text{ mol }P_4O_{10}}{1\ \cancel{mol\ P_4}} = 0.202\text{ mol }P_4O_{10}$$

Calculate the moles of product (P_4O_{10}) formed.

To calculate the mass of P_4O_{10}, multiply moles of P_4O_{10} by the conversion factor that relates mass and moles—molar mass.

$$0.202\ \cancel{mol\ P_4O_{10}} \times \frac{283.9\text{ g }P_4O_{10}}{1\ \cancel{mol\ P_4O_{10}}} = \mathbf{57.3\text{ g }P_4O_{10}}$$

Calculate the mass of the product P_4O_{10}.

Because O_2 is in excess, only part of the available O_2 is consumed. Use the limiting reactant, P_4, to determine the moles and mass of O_2 used.

$$0.202\ \cancel{\text{mol } P_4} \times \frac{5\ \text{mol } O_2}{1\ \cancel{\text{mol } P_4}} = 1.01\ \text{mol } O_2$$

Multiply the moles of limiting reactant by the mole ratio to determine moles of excess reactant needed.

Convert moles of O_2 consumed to mass of O_2 consumed.

$$1.01\ \cancel{\text{mol } O_2} \times \frac{32.00\ \text{g } O_2}{1\ \cancel{\text{mol } O_2}} = 32.3\ \text{g } O_2$$

Multiply the moles of O_2 by the molar mass.

Calculate the amount of excess O_2.

50.0 g O_2 available − 32.3 g O_2 consumed = **17.7 g O_2 in excess**

Subtract the mass of O_2 used from the mass available.

3 Evaluate the Answer

All values have a minimum of three significant figures, so the mass of P_4O_{10} is correctly stated with three digits. The mass of excess O_2 (17.7 g) is found by subtracting two numbers that are accurate to the first decimal place. Therefore, the mass of excess O_2 correctly shows one decimal place. The sum of the O_2 that was consumed (32.3 g) and the given mass of P_4 (25.0 g) is 57.3 g, the calculated mass of the product P_4O_{10}.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

23. The reaction between solid sodium and iron(III) oxide is one in a series of reactions that inflates an automobile airbag: $6Na(s) + Fe_2O_3(s) \longrightarrow 3Na_2O(s) + 2Fe(s)$. If 100.0 g of Na and 100.0 g of Fe_2O_3 are used in this reaction, determine the following.

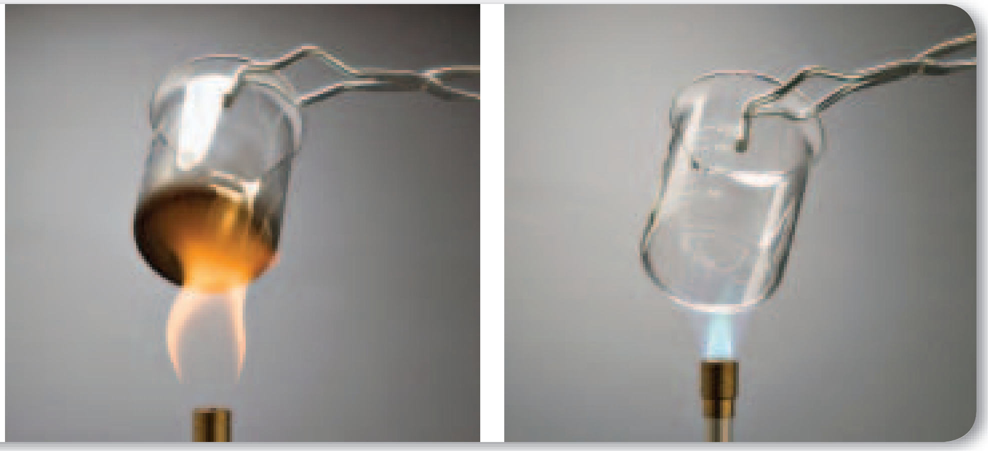
- **a.** limiting reactant
- **b.** reactant in excess
- **c.** mass of solid iron produced
- **d.** mass of excess reactant that remains after the reaction is complete

24. Challenge Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. A plant has 88.0 g of carbon dioxide and 64.0 g of water available for photosynthesis.

- **a.** Write the balanced chemical equation for the reaction.
- **b.** Determine the limiting reactant.
- **c.** Determine the excess reactant.
- **d.** Determine the mass in excess.
- **e.** Determine the mass of glucose produced.

Connection to Biology Your body needs vitamins, minerals, and elements in small amounts to facilitate normal metabolic reactions. A lack of these substances can lead to abnormalities in growth, development, and the functioning of your body's cells. Phosphorus, for example, is an essential element in living systems; phosphate groups occur regularly in strands of DNA. Potassium is needed for proper nerve function, muscle control, and blood pressure. A diet low in potassium and high in sodium might be a factor in high blood pressure. Another example is vitamin B-12. Without adequate vitamin B-12, the body is unable to synthesize DNA properly, affecting the production of red blood cells.

■ **Figure 11.7** With insufficient oxygen, the burner on the left burns with a yellow, sooty flame. The burner on the right burns hot and clean because an excess of oxygen is available to react completely with the methane gas.



Why use an excess of a reactant? Many reactions stop while portions of the reactants are still present in the reaction mixture. Because this is inefficient and wasteful, chemists have found that by using an excess of one reactant—often the least expensive one—reactions can be driven to continue until all of the limiting reactant is used up. Using an excess of one reactant can also speed up a reaction.

Figure 11.7 shows an example of how controlling the amount of a reactant can increase efficiency. Your lab likely uses the type of Bunsen burner shown in the figure. If so, you know that this type of burner has a control that lets you adjust the amount of air that mixes with the methane gas. How efficiently the burner operates depends on the ratio of oxygen to methane gas in the fuel mixture. When the air is limited, the resulting flame is yellow because of glowing bits of unburned fuel. This unburned fuel leaves soot (carbon) deposits on glassware. Fuel is wasted because the amount of energy released is less than the amount that could have been produced if enough oxygen were available. When sufficient oxygen is present in the combustion mixture, the burner produces a hot, intense blue flame. No soot is deposited because the fuel is completely converted to carbon dioxide and water vapor.

Section 11.3 Assessment

Section Summary

- The limiting reactant is the reactant that is completely consumed during a chemical reaction. Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the limiting reactant.

25. MAIN Idea **Describe** the reason why a reaction between two substances comes to an end.

26. **Identify** the limiting and the excess reactant in each reaction.
    a. Wood burns in a campfire.
    b. Airborne sulfur reacts with the silver plating on a teapot to produce tarnish (silver sulfide).
    c. Baking powder in batter decomposes to produce carbon dioxide.

27. **Analyze** Tetraphosphorus trisulphide (P_4S_3) is used in the match heads of some matches. It is produced in the reaction $8P_4 + 3S_8 \longrightarrow 8P_4S_3$. Determine which of the following statements are incorrect, and rewrite the incorrect statements to make them correct.
    a. 4 mol P_4 reacts with 1.5 mol S_8 to form 4 mol P_4S_3.
    b. Sulfur is the limiting reactant when 4 mol P_4 and 4 mol S_8 react.
    c. 6 mol P_4 reacts with 6 mol S_8, forming 1320 g P_4S_3.

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Section 11.4

Objectives

- **Calculate** the theoretical yield of a chemical reaction from data.
- **Determine** the percent yield for a chemical reaction.

Review Vocabulary

process: a series of actions or operations

New Vocabulary

theoretical yield
actual yield
percent yield

Percent Yield

MAIN Idea Percent yield is a measure of the efficiency of a chemical reaction.

Real-World Reading Link Imagine that you are practicing free throws and you take 100 practice shots. Theoretically, you could make all 100 shots. In actuality, however, you know you will not make all of the shots. Chemical reactions also have theoretical and actual outcomes.

How much product?

While solving stoichiometric problems in this chapter, you might have concluded that chemical reactions always proceed in the laboratory according to the balanced equation and produce the calculated amount of product. This, however, is not the case. Just as you are unlikely to make 100 out of 100 free throws during basketball practice, most reactions never succeed in producing the predicted amount of product. Reactions do not go to completion or yield as expected for a variety of reasons. Liquid reactants and products might adhere to the surfaces of their containers or evaporate. In some instances, products other than the intended ones might be formed by competing reactions, thus reducing the yield of the desired product. Or, as shown in **Figure 11.8,** some amount of any solid product is usually left behind on filter paper or lost in the purification process. Because of these problems, chemists need to know how to gauge the yield of a chemical reaction.

Theoretical and Actual Yields In many of the stoichiometric calculations you have performed, you have calculated the amount of product produced from a given amount of reactant. The answer you obtained is the theoretical yield of the reaction. The **theoretical yield** is the maximum amount of product that can be produced from a given amount of reactant.

A chemical reaction rarely produces the theoretical yield of product. A chemist determines the actual yield of a reaction through a careful experiment in which the mass of the product is measured. The **actual yield** is the amount of product produced when the chemical reaction is carried out in an experiment.

Figure 11.8 Silver chromate is formed when potassium chromate is added to silver nitrate. Note that some of the precipitate is left behind on filter paper. Still more of the precipitate is lost because it adheres to the sides of the beaker.

Percent yield Chemists need to know how efficient a reaction is in producing the desired product. One way of measuring efficiency is by means of percent yield. **Percent yield** of product is the ratio of the actual yield to the theoretical yield expressed as a percent.

Percent Yield

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

The actual yield divided by the theoretical yield multiplied by 100 is the percent yield.

EXAMPLE Problem 11.6

Math Handbook
Percents
page 965

Percent Yield Solid silver chromate (Ag_2CrO_4) forms when potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g of silver nitrate ($AgNO_3$). Determine the theoretical yield of Ag_2CrO_4. Calculate the percent yield if the reaction yields 0.455 g of Ag_2CrO_4.

1 Analyze the Problem

You know the mass of a reactant and the actual yield of the product. Write the balanced chemical equation, and calculate theoretical yield by converting grams of $AgNO_3$ to moles of $AgNO_3$, moles of $AgNO_3$ to moles of Ag_2CrO_4, and moles of Ag_2CrO_4 to grams of Ag_2CrO_4. Calculate the percent yield from the actual yield and the theoretical yield.

Known

mass of silver nitrate = 0.500 g $AgNO_3$

actual yield = 0.455 g Ag_2CrO_4

Unknown

theoretical yield = ? g Ag_2CrO_4

percent yield = ? % Ag_2CrO_4

2 Solve for the Unknown

$$\overset{0.500\text{ g}}{2AgNO_3(aq)} + K_2CrO_4(aq) \rightarrow \overset{?\text{ g}}{Ag_2CrO_4(s)} + 2KNO_3(aq)$$

Write the balanced chemical equation, and identify the known and the unknown.

$$0.500\ \cancel{\text{g } AgNO_3} \times \frac{1 \text{ mol } AgNO_3}{169.9\ \cancel{\text{g } AgNO_3}} = 2.94 \times 10^{-3} \text{ mol } AgNO_3$$

Use molar mass to convert grams of $AgNO_3$ to moles of $AgNO_3$.

$$2.94 \times 10^{-3}\ \cancel{\text{mol } AgNO_3} \times \frac{1 \text{ mol } Ag_2CrO_4}{2\ \cancel{\text{mol } AgNO_3}} = 1.47 \times 10^{-3} \text{ mol } Ag_2CrO_4$$

Use the mole ratio to convert moles of $AgNO_3$ to moles of Ag_2CrO_4.

$$1.47 \times 10^{-3}\ \cancel{\text{mol } Ag_2CrO_4} \times \frac{331.7 \text{ g } Ag_2CrO_4}{1\ \cancel{\text{mol } Ag_2CrO_4}} = \mathbf{0.488 \text{ g } Ag_2CrO_4}$$

Calculate the theoretical yield.

$$\frac{0.455\ \cancel{\text{g } Ag_2CrO_4}}{0.488\ \cancel{\text{g } Ag_2CrO_4}} \times 100 = \mathbf{93.2\% \ Ag_2CrO_4}$$

Calculate the percent yield.

3 Evaluate the Answer

The quantity with the fewest significant figures has three, so the percent is correctly stated with three digits. The molar mass of Ag_2CrO_4 is about twice the molar mass of $AgNO_3$, and the ratio of moles of $AgNO_3$ to moles of Ag_2CrO_4 in the equation is 2:1. Therefore, 0.500 g of $AgNO_3$ should produce about the same mass of Ag_2CrO_4. The actual yield of Ag_2CrO_4 is close to 0.500 g, so a percent yield of 93.2% is reasonable.

PRACTICE Problems

Extra Practice Page 983 and glencoe.com

28. Aluminum hydroxide ($Al(OH)_3$) is often present in antacids to neutralize stomach acid (HCl). The reaction occurs as follows: $Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$. If 14.0 g of $Al(OH)_3$ is present in an antacid tablet, determine the theoretical yield of $AlCl_3$ produced when the tablet reacts with HCl.

29. Zinc reacts with iodine in a synthesis reaction: $Zn + I_2 \longrightarrow ZnI_2$.

a. Determine the theoretical yield if 1.912 mol of zinc is used.

b. Determine the percent yield if 515.6 g of product is recovered.

30. **Challenge** When copper wire is placed into a silver nitrate solution ($AgNO_3$), silver crystals and copper(II) nitrate ($Cu(NO_3)_2$) solution form.

a. Write the balanced chemical equation for the reaction.

b. If a 20.0-g sample of copper is used, determine the theoretical yield of silver.

c. If 60.0 g of silver is recovered from the reaction, determine the percent yield of the reaction.

DATA ANALYSIS LAB

Based on Real Data[1, 2]

Analyze and Conclude

Can rocks on the Moon provide an effective oxygen source for future lunar missions? Although the Moon has no atmosphere and thus no oxygen, its surface is covered with rocks and soil made from oxides. Scientists, looking for an oxygen source for future long-duration lunar missions, are researching ways to extract oxygen from lunar soil and rock. Analysis of samples collected during previous lunar missions provided scientists with the data shown in the table. The table identifies the oxides in lunar soil as well as each oxide's percent-by-weight of the soil.

Data and Observations

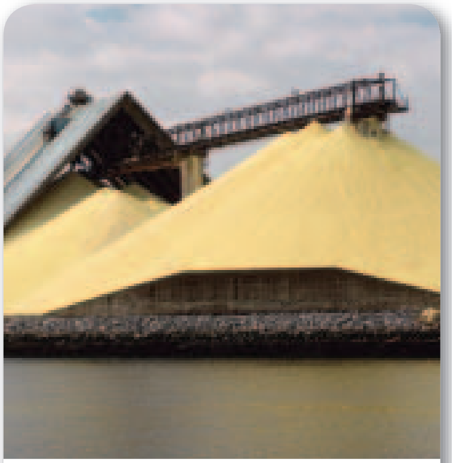
Moon-Rock Data[1]

Oxide	% Weight of Soil
SiO_2	47.3%
Al_2O_3	17.8%
CaO	11.4%
FeO	10.5%
MgO	9.6%
TiO_2	1.6%
Na_2O	0.7%
K_2O	0.6%
Cr_2O_3	0.2%
MnO	0.1%

[1]Data obtained from: McKay, et al. 1994. JSC-1: A new lunar soil stimulant. *Engineering, Construction, and Operations in Space* IV: 857–866, American Society of Civil Engineers.
[2]Data obtained from: Berggren, et al. 2005. Carbon monoxide silicate reduction system. *Space Resources Roundtable* VII.

Think Critically

1. **Calculate** For each of the oxides listed in the table, determine the mass (in grams) that would exist in 1.00 kg of lunar soil.
2. **Apply** Scientists want to release the oxygen from its metal oxide using a decomposition reaction: metal oxide $\longrightarrow$ metal + oxygen. To assess the viability of this idea, determine the amount of oxygen per kilogram contained in each of the oxides found in lunar soil.
3. **Identify** What oxide would yield the most oxygen per kilogram? The least?
4. **Determine** the theoretical yield of oxygen from the oxides present in a 1.00-kg sample of lunar soil.
5. **Calculate** Using methods currently available, scientists can produce 15 kg of oxygen from 100 kg of lunar soil. What is the percent yield of the process?



■ **Figure 11.9** Sulfur, such as these piles at Vancouver Harbor, can be extracted from petroleum products by a chemical process. Sulfur is also mined by forcing hot water into underground deposits and pumping the liquid sulfur to the surface.

Percent Yield in the Marketplace

Percent yield is important in the cost effectiveness of many industrial manufacturing processes. For example, the sulfur shown in **Figure 11.9** is used to make sulfuric acid (H_2SO_4). Sulfuric acid is an important chemical because it is a raw material used to make products such as fertilizers, detergents, pigments, and textiles. The cost of sulfuric acid affects the cost of many of the consumer items you use every day. The first two steps in the manufacturing process are shown below.

Step 1 $$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$

Step 2 $$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

In the final step, SO_3 combines with water to produce H_2SO_4.

Step 3 $$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)$$

The first step, the combustion of sulfur, produces an almost 100% yield. The second step also produces a high yield if a catalyst is used at the relatively low temperature of 400°C. A catalyst is a substance that speeds a reaction but does not appear in the chemical equation. Under these conditions, the reaction is slow. Raising the temperature increases the reaction rate but decreases the yield.

To maximize yield and minimize time in the second step, engineers have devised a system in which the reactants, O_2 and SO_2, are passed over a catalyst at 400°C. Because the reaction releases a great deal of heat, the temperature gradually increases with an accompanying decrease in yield. Thus, when the temperature reaches approximately 600°C, the mixture is cooled and then passed over the catalyst again. A total of four passes over the catalyst with cooling between passes results in a yield greater than 98%.

Section 11.4 Assessment

Section Summary

- The theoretical yield of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The actual yield is the amount of product produced. Actual yield must be obtained through experimentation.
- Percent yield is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.

31. MAIN Idea **Identify** which type of yield—theoretical yield, actual yield, or percent yield—is a measure of the efficiency of a chemical reaction.

32. List several reasons why the actual yield from a chemical reaction is not usually equal to the theoretical yield.

33. Explain how percent yield is calculated.

34. Apply In an experiment, you combine 83.77 g of iron with an excess of sulfur and then heat the mixture to obtain iron(III) sulfide.

$$2Fe(s) + 3S(s) \longrightarrow Fe_2S_3(s)$$

What is the theoretical yield, in grams, of iron(III) sulfide?

35. Calculate the percent yield of the reaction of magnesium with excess oxygen: $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Reaction Data	
Mass of empty crucible	35.67 g
Mass of crucible and Mg	38.06 g
Mass of crucible and MgO (after heating)	39.15 g

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Battling Resistant Strains

The human immunodeficiency virus (HIV), the virus that causes AIDS, has proven to be among the most incurable foes ever faced by modern medical science. One reason for this is the virus's remarkable ability to adapt. Resistant strains of the virus appear quickly, rendering obsolete the newest and most powerful AIDS drugs. Now some researchers are using the virus's adaptability as a way to fight it.

Selecting resistance PA-457 is a promising new anti-HIV drug synthesized from betulinic acid, an organic compound derived from some plants, including the bark of birch trees. To find out just what PA-457 does to HIV, known as the drug's mechanism of action, researchers took what might seem a strange step: they encouraged samples of HIV to develop resistance to PA-457.

Researchers subjected HIV samples to small doses of PA-457. Using a low dose made it more likely that some of the virus would survive the treatment and possibly develop resistance. Those viruses that survived exposure were collected, and their genetic sequences were examined. The surviving viruses were found to have a mutation in the genes that control how the virus builds a structure called a capsid, shown in **Figure 1.**

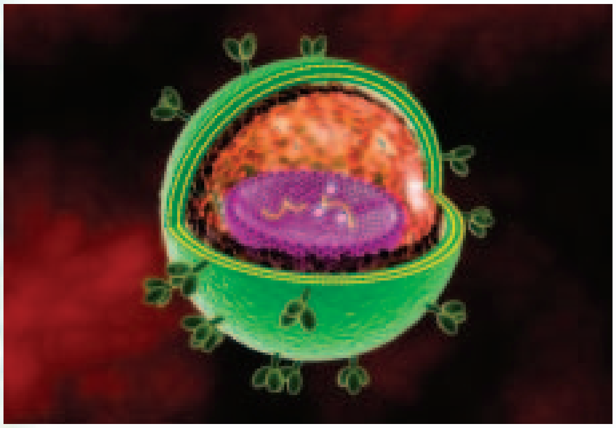


Figure 1 In a normal HIV virus, the capsid forms a protective coating around the genetic material.

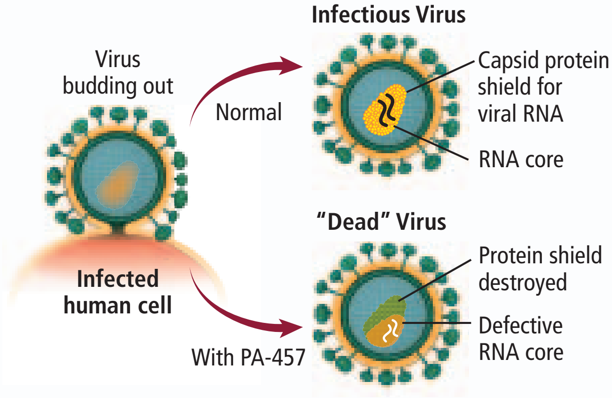


Figure 2 When treated with PA-457, the HIV capsid becomes misshapen and collapses, resulting in the death of the virus.

Surprise attack This finding was surprising, because it showed that, unlike most drugs, PA-457 attacks the HIV structure, rather than the enzymes that help HIV reproduce, as illustrated in **Figure 2.** This makes PA-457 among the first of a new class of HIV drugs known as maturation inhibitors—drugs that can prevent the virus from maturing during the late stages in its development.

Slowing evolution The hope is that because PA-457 and other maturation inhibitors attack the HIV structure, resistance will be slower to develop. Even so, maturation inhibitors will likely be prescribed in combination with other AIDS drugs that attack HIV at different stages of its life cycle.

This practice, called multidrug therapy, makes it harder for HIV to develop resistance because any surviving virus would need to have multiple mutations—at least one for each anti-HIV drug. These mutations are less likely to occur at the same time.

WRITING in Chemistry

Research how scientists determine the safe dosing level for an experimental drug. Discuss how a drug's effectiveness must be balanced with its potential toxicity and side effects. For more information on how a therapeutic dose is determined, visit glencoe.com.

CHEMLAB

DETERMINE THE MOLE RATIO

Chemistry Online

Probeware Alternate CBL instructions can be found at glencoe.com.

Background: **Iron reacts with copper(II) sulfate ($CuSO_4$). By measuring the mass of iron that reacts and the mass of copper metal produced, you can calculate the experimental mole ratio.**

Question: *How does the experimental mole ratio compare with the theoretical mole ratio?*

Materials

copper(II) sulfate pentahydrate ($CuSO_4$•$5H_2O$)
iron metal filings (20 mesh)
distilled water
150-mL beaker
100-mL graduated cylinder
hot plate
beaker tongs
balance
stirring rod
400-mL beaker
weighing paper

Safety Precautions

WARNING: ***Hot plates can cause burns. Turn off hot plates when not in use. Use only GFCI-protected circuits.***

Procedure

1. Read and complete the lab safety form.
2. Measure the mass of a clean, dry 150-mL beaker. Record all measurements in a data table.
3. Place approximately 12 g $CuSO_4$•$5H_2O$ into the 150-mL beaker, and measure the combined mass.
4. Add 50 mL of distilled water to the $CuSO_4$•$5H_2O$. Place the mixture on a hot plate set at medium, and stir until all of the solid dissolves (do not boil). Using tongs, remove the beaker from the hot plate.
5. Measure about 2 g of iron filings onto a piece of weighing paper. Measure the mass of the filings.
6. While stirring, slowly add the iron filings to the hot copper(II) sulfate solution. Be careful not to splash the hot solution.
7. Allow the reaction mixture to sit for 5 min.
8. Use the stirring rod to decant (pour off) the liquid into a 400-mL beaker. Be careful to decant only the liquid—leave the solid copper metal behind.
9. Add 15 mL of distilled water to the copper solid, and carefully swirl the beaker to wash the copper. Decant the liquid into the 400-mL beaker.
10. Repeat Step 9 two more times.
11. Place the beaker containing the wet copper on the hot plate. Use low heat to dry the copper.
12. After the copper is dry, use tongs to remove the beaker from the hot plate and allow it to cool.
13. Measure the mass of the beaker and the copper.
14. **Cleanup and Disposal** The dry copper can be placed in a waste container. Moisten any residue that sticks to the beaker, and wipe it out using a paper towel. Pour the unreacted copper(II) sulfate and iron(II) sulfate solutions into a large beaker. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Apply** Write a balanced chemical equation for the reaction and calculate the mass of copper (Cu) that should have formed from the sample of iron (Fe) used. This mass is the theoretical yield.
2. **Interpret Data** Using your data, determine the mass and the moles of copper produced. Calculate the moles of iron used, and determine the whole-number iron-to-copper mole ratio and percent yield.
3. **Compare and Contrast** Compare the theoretical iron-to-copper mole ratio to the mole ratio you calculated using the experimental data.
4. **Error Analysis** Identify sources of the error that resulted in deviation from the mole ratio given in the balanced chemical equation.

INQUIRY EXTENSION

Compare your results with those of several other lab teams. Create a hypothesis to explain any differences.

Matt Meadows

CHAPTER 11 Study Guide

BIG Idea Mass relationships in chemical reactions confirm the law of conservation of mass.

Section 11.1 Defining Stoichiometry

MAIN Idea The amount of each reactant present at the start of a chemical reaction determines how much product can form.

Vocabulary
- mole ratio (p. 371)
- stoichiometry (p. 368)

Key Concepts
- Balanced chemical equations can be interpreted in terms of moles, mass, and representative particles (atoms, molecules, formula units).
- The law of conservation of mass applies to all chemical reactions.
- Mole ratios are derived from the coefficients of a balanced chemical equation. Each mole ratio relates the number of moles of one reactant or product to the number of moles of another reactant or product in the chemical reaction.

Section 11.2 Stoichiometric Calculations

MAIN Idea The solution to every stoichiometric problem requires a balanced chemical equation.

Key Concepts
- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.
- Stoichiometric problems make use of mole ratios to convert between mass and moles.

Section 11.3 Limiting Reactants

MAIN Idea A chemical reaction stops when one of the reactants is used up.

Vocabulary
- excess reactant (p. 379)
- limiting reactant (p. 379)

Key Concepts
- The limiting reactant is the reactant that is completely consumed during a chemical reaction. Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the limiting reactant.

Section 11.4 Percent Yield

MAIN Idea Percent yield is a measure of the efficiency of a chemical reaction.

Vocabulary
- actual yield (p. 385)
- percent yield (p. 386)
- theoretical yield (p. 385)

Key Concepts
- The theoretical yield of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The actual yield is the amount of product produced. Actual yield must be obtained through experimentation.
- Percent yield is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.

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

CHAPTER 11 Assessment

Section 11.1

Mastering Concepts

36. Why must a chemical equation be balanced before you can determine mole ratios?

37. What relationships can be determined from a balanced chemical equation?

38. Explain why mole ratios are central to stoichiometric calculations.

39. What is the mole ratio that can convert from moles of A to moles of B?

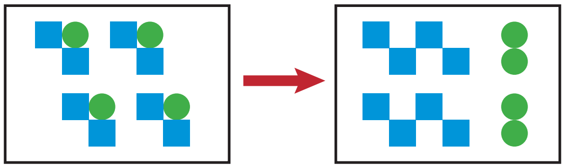
40. Why are coefficients used in mole ratios instead of subscripts?

41. Explain how the conservation of mass allows you to interpret a balanced chemical equation in terms of mass.

42. When heated by a flame, ammonium dichromate decomposes, producing nitrogen gas, solid chromium(III) oxide, and water vapor.

$$(NH_4)2Cr_2O_7 \longrightarrow N_2 + Cr_2O_3 + 4H_2O$$

Write the mole ratios for this reaction that relate ammonium dichromate to the products.



■ **Figure 11.10**

43. **Figure 11.10** depicts an equation with squares representing Element M and circles representing Element N. Write a balanced equation to represent the picture shown, using smallest whole-number ratios. Write mole ratios for this equation.

Mastering Problems

44. Interpret the following equation in terms of particles, moles, and mass.

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

45. **Smelting** When tin(IV) oxide is heated with carbon in a process called smelting, the element tin can be extracted.

$$SnO_2(s) + 2C(s) \longrightarrow Sn(l) + 2CO(g)$$

Interpret the chemical equation in terms of particles, moles, and mass.

46. When solid copper is added to nitric acid, copper(II) nitrate, nitrogen dioxide, and water are produced. Write the balanced chemical equation for the reaction. List six mole ratios for the reaction.

47. When hydrochloric acid solution reacts with lead(II) nitrate solution, lead(II) chloride precipitates and a solution of nitric acid is produced.

a. Write the balanced chemical equation for the reaction.

b. Interpret the equation in terms of molecules and formula units, moles, and mass.

48. When aluminum is mixed with iron(III) oxide, iron metal and aluminum oxide are produced, along with a large quantity of heat. What mole ratio would you use to determine moles of Fe if moles of Fe_2O_3 is known?

$$Fe_2O_3(s) + 2Al(s) \longrightarrow 2Fe(s) + Al_2O_3(s) + \text{heat}$$

49. Solid silicon dioxide, often called silica, reacts with hydrofluoric acid (HF) solution to produce the gas silicon tetrafluoride and water.

a. Write the balanced chemical equation for the reaction.

b. List three mole ratios, and explain how you would use them in stoichiometric calculations.

50. **Chrome** The most important commercial ore of chromium is chromite ($FeCr_2O_4$). One of the steps in the process used to extract chromium from the ore is the reaction of chromite with coke (carbon) to produce ferrochrome ($FeCr_2$).

$$2C(s) + FeCr_2O_4(s) \longrightarrow FeCr_2(s) + 2CO_2(g)$$

What mole ratio would you use to convert from moles of chromite to moles of ferrochrome?

51. **Air Pollution** The pollutant SO_2 is removed from the air by in a reaction that also involves calcium carbonate and oxygen. The products of this reaction are calcium sulfate and carbon dioxide. Determine the mole ratio you would use to convert moles of SO_2 to moles of $CaSO_4$.

52. Two substances, W and X, react to form the products Y and Z. **Table 11.2** shows the moles of the reactants and products involved when the reaction was carried out. Use the data to determine the coefficients that will balance the equation $W + X \longrightarrow Y + Z$.

Table 11.2 Reaction Data

Moles of Reactants		Moles of Products	
W	X	Y	Z
0.90	0.30	0.60	1.20

53. **Antacids** Magnesium hydroxide is an ingredient in some antacids. Antacids react with excess hydrochloric acid in the stomach to relieve indigestion.

$$\_\_\_Mg(OH)_2 + \_\_\_HCl \longrightarrow \_\_\_ MgCl_2 + \_\_\_H_2O$$

a. Balance the reaction of $Mg(OH)_2$ with HCl.

b. Write the mole ratio that would be used to determine the number of moles of $MgCl_2$ produced when HCl reacts with $Mg(OH)_2$.

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Section 11.2

Mastering Concepts

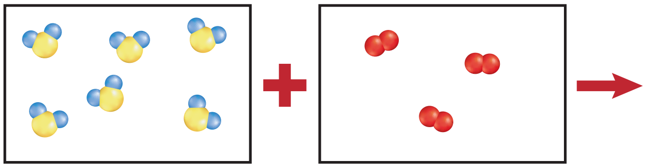
54. What is the first step in all stoichiometric calculations?

55. What information does a balanced equation provide?

56. On what law is stoichometry based, and how do the calculations support this law?

57. How is molar mass used in some stoichiometric calculations?

58. What information must you have in order to calculate the mass of product formed in a chemical reaction?



■ **Figure 11.11**

59. Each box in **Figure 11.11** represents the contents of a flask. One flask contains hydrogen sulfide, and the other contains oxygen. When the contents of the flasks are mixed, a reaction occurs and water vapor and sulfur are produced. In the figure, the red circles represent oxygen, the yellow circles represent sulfur, and blue circles represent hydrogen.
a. Write the balanced chemical equation for the reaction.
b. Using the same color code, sketch a representation of the flask after the reaction occurs.

Mastering Problems

60. **Ethanol** (C_2H_5OH), also known as grain alcohol, can be made from the fermentation of sugar ($C_6H_{12}O_6$). The unbalanced chemical equation for the reaction is shown below.

$$\_\_\_C_6H_{12}O_6 \longrightarrow \_\_\_C_2H_5OH + \_\_\_CO_2$$

Balance the chemical equation and determine the mass of C_2H_5OH produced from 750 g of $C_6H_{12}O_6$.

61. **Welding** If 5.50 mol of calcium carbide (CaC_2) reacts with an excess of water, how many moles of acetylene (C_2H_2), a gas used in welding, will be produced?

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

62. **Antacid Fizz** When an antacid tablet dissolves in water, the fizz is due to a reaction between sodium hydrogen carbonate ($NaHCO_3$), also called sodium bicarbonate, and citric acid ($H_3C_6H_5O_7$).

$$3NaHCO_3(aq) + H_3C_6H_5O_7(aq) \longrightarrow 3CO_2(g) + 3H_2O(l) + Na_3C_6H_5O_7(aq)$$

How many moles of $Na_3C_6H_5O_7$ can be produced if one tablet containing 0.0119 mol of $NaHCO_3$ is dissolved?

63. **Esterification** The process in which an organic acid and an alcohol react to form an ester and water is known as esterification. Ethyl butanoate ($C_3H_7COOC_2H_5$), an ester, is formed when the alcohol ethanol (C_2H_5OH) and butanoic acid (C_3H_7COOH) are heated in the presence of sulfuric acid.

$$C_2H_5OH(l) + C_3H_7COOH(l) \longrightarrow C_3H_7COOC_2H_5(l) + H_2O(l)$$

Determine the mass of ethyl butanoate produced if 4.50 mol of ethanol is used.

64. **Greenhouse Gas** Carbon dioxide is a greenhouse gas that is linked to global warming. It is released into the atmosphere through the combustion of octane (C_8H_{18}) in gasoline. Write the balanced chemical equation for the combustion of octane and calculate the mass of octane needed to release 5.00 mol of CO_2.

65. A solution of potassium chromate reacts with a solution of lead(II) nitrate to produce a yellow precipitate of lead(II) chromate and a solution of potassium nitrate.
a. Write the balanced chemical equation.
b. Starting with 0.250 mol of potassium chromate, determine the mass of lead chromate formed.

66. **Rocket Fuel** The exothermic reaction between liquid hydrazine (N_2H_2) and liquid hydrogen peroxide (H_2O_2) is used to fuel rockets. The products of this reaction are nitrogen gas and water.
a. Write the balanced chemical equation.
b. How much hydrazine, in grams, is needed to produce 10.0 mol of nitrogen gas?

67. Chloroform ($CHCl_3$), an important solvent, is produced by a reaction between methane and chlorine.

$$CH_4(g) + 3Cl_2(g) \longrightarrow CHCl_3(g) + 3HCl(g)$$

How much CH_4, in grams, is needed to produce 50.0 grams of $CHCl_3$?

68. **Oxygen Production** The Russian Space Agency uses potassium superoxide (KO_2) for the chemical oxygen generators in their space suits.

$$4KO_2 + 2H_2O + 4CO_2 \longrightarrow 4KHCO_3 + 3O_2$$

Complete **Table 11.3.**

Table 11.3 Oxygen Generation Reaction Data

Mass KO_2	Mass H_2O	Mass CO_2	Mass $KHCO_3$	Mass O_2
				380 g

69. **Gasohol** is a mixture of ethanol and gasoline. Balance the equation, and determine the mass of CO_2 produced from the combustion of 100.0 g of ethanol.

$$C_2H_5OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

70. **Car Battery** Car batteries use lead, lead(IV) oxide, and a sulfuric acid solution to produce an electric current. The products of the reaction are lead(II) sulfate in solution and water.
  a. Write the balanced equation for the reaction.
  b. Determine the mass of lead(II) sulfate produced when 25.0 g of lead reacts with an excess of lead(IV) oxide and sulfuric acid.

71. To extract gold from its ore, the ore is treated with sodium cyanide solution in the presence of oxygen and water.

$$4Au(s) + 8NaCN(aq) + O_2(g) + 2H_2O(l) \longrightarrow 4NaAu(CN)_2(aq) + 4NaOH(aq)$$

  a. Determine the mass of gold that can be extracted if 25.0 g of sodium cyanide is used.
  b. If the mass of the ore from which the gold was extracted is 150.0 g, what percentage of the ore is gold?

72. **Film** Photographic film contains silver bromide in gelatin. Once exposed, some of the silver bromide decomposes, producing fine grains of silver. The unexposed silver bromide is removed by treating the film with sodium thiosulfate. Soluble sodium silver thiosulfate ($Na_3Ag(S_2O_3)_2$) is produced.

$$AgBr(s) + 2Na_2S_2O_3(aq) \longrightarrow Na_3Ag(S_2O_3)_2(aq) + NaBr(aq)$$

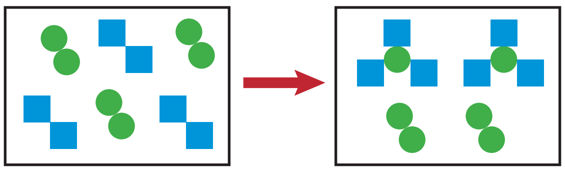
Determine the mass of $Na_3Ag(S_2O_3)_2$ produced if 0.275 g of AgBr is removed.

Section 11.3

Mastering Concepts

73. How is a mole ratio used to find the limiting reactant?

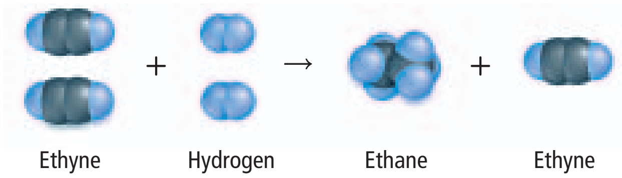
74. Explain why the statement, "The limiting reactant is the reactant with the lowest mass" is incorrect.



■ **Figure 11.12**

75. **Figure 11.12** uses squares to represent Element M and circles to represent Element N.
  a. Write the balanced equation for the reaction.
  b. If each square represents 1 mol of M and each circle represents 1 mol of N, how many moles of M and N were present at the start of the reaction?
  c. How many moles of product form? How many moles of Element M and Element N are unreacted?
  d. Identify the limiting reactant and the excess reactant.

Mastering Problems



■ **Figure 11.13**

76. The reaction between ethyne (C_2H_2) and hydrogen (H_2) is illustrated in **Figure 11.13.** The product is ethane (C_2H_6). Which is the limiting reactant? Which is the excess reactant? Explain.

77. **Nickel-Iron Battery** In 1901, Thomas Edison invented the nickel-iron battery. The following reaction takes place in the battery.

$$Fe(s) + 2NiO(OH)(s) + 2H_2O(l) \longrightarrow Fe(OH)_2(s) + 2Ni(OH)_2(aq)$$

How many mol of $Fe(OH)_2$ is produced when 5.00 mol of Fe and 8.00 mol of NiO(OH) react?

78. One of the few xenon compounds that form is cesium xenon heptafluoride ($CsXeF_7$). How many moles of $CsXeF_7$ can be produced from the reaction of 12.5 mol of cesium fluoride with 10.0 mol of xenon hexafluoride?

$$CsF(s) + XeF_6(s) \longrightarrow CsXeF_7(s)$$

79. **Iron Production** Iron is obtained commercially by the reaction of hematite (Fe_2O_3) with carbon monoxide. How many grams of iron is produced when 25.0 mol of hematite reacts with 30.0 mol of carbon monoxide?

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

80. The reaction of chlorine gas with solid phosphorus (P_4) produces solid phosphorus pentachloride. When 16.0 g of chlorine reacts with 23.0 g of P_4, which reactant is limiting? Which reactant is in excess?

81. **Alkaline Battery** An alkaline battery produces electrical energy according to this equation.

$$Zn(s) + 2MnO_2(s) + H_2O(l) \longrightarrow Zn(OH)_2(s) + Mn_2O_3(s)$$

  a. Determine the limiting reactant if 25.0 g of Zn and 30.0 g of MnO_2 are used.
  b. Determine the mass of $Zn(OH)_2$ produced.

82. Lithium reacts spontaneously with bromine to produce lithium bromide. Write the balanced chemical equation for the reaction. If 25.0 g of lithium and 25.0 g of bromine are present at the beginning of the reaction, determine
  a. the limiting reactant.
  b. the mass of lithium bromide produced.
  c. the excess reactant and the excess mass.

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Section 11.4

Mastering Concepts

83. What is the difference between actual yield and theoretical yield?

84. How are actual yield and theoretical yield determined?

85. Can the percent yield of a chemical reaction be more than 100%? Explain your answer.

86. What relationship is used to determine the percent yield of a chemical reaction?

87. What experimental information do you need in order to calculate both the theoretical and the percent yield of any chemical reaction?

88. A metal oxide reacts with water to produce a metal hydroxide. What additional information would you need to determine the percent yield of metal hydroxide from this reaction?

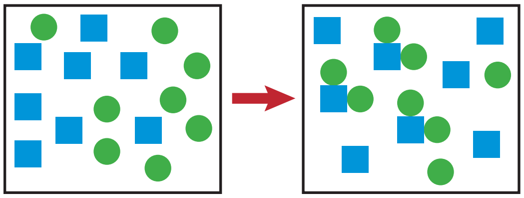

Figure 11.14

89. Examine the reaction represented in **Figure 11.14.** Determine if the reaction went to completion. Explain your answer, and calculate the percent yield of the reaction.

Mastering Problems

90. Ethanol (C_2H_5OH) is produced from the fermentation of sucrose ($C_{12}H_{22}O_{11}$) in the presence of enzymes.

$$C_{12}H_{22}O_{11}(aq) + H_2O(g) \longrightarrow 4C_2H_5OH(l) + 4CO_2(g)$$

Determine the theoretical yield and the percent yield of ethanol if 684 g of sucrose undergoes fermentation and 349 g of ethanol is obtained.

91. Lead(II) oxide is obtained by roasting galena, lead(II) sulfide, in air. The unbalanced equation is:

$$PbS(s) + O_2(g) \longrightarrow PbO(s) + SO_2(g)$$

a. Balance the equation, and determine the theoretical yield of PbO if 200.0 g of PbS is heated.
b. What is the percent yield if 170.0 g of PbO is obtained?

92. Upon heating, calcium carbonate ($CaCO_3$) decomposes to calcium oxide (CaO) and carbon dioxide (CO_2).
a. Determine the theoretical yield of CO_2 if 235.0 g of $CaCO_3$ is heated.
b. What is the percent yield of CO_2 if 97.5 g of CO_2 is collected?

93. Hydrofluoric acid solutions cannot be stored in glass containers because HF reacts readily with silica dioxide in glass to produce hexafluorosilicic acid (H_2SiF_6).

$$SiO_2(s) + 6HF(aq) \longrightarrow H_2SiF_6(aq) + 2H_2O(l)$$

40.0 g SiO_2 and 40.0 g HF react to yield 45.8 g H_2SiF_6.
a. What is the limiting reactant?
b. What is the mass of the excess reactant?
c. What is the theoretical yield of H_2SiF_6?
d. What is the percent yield?

94. Van Arkel Process Pure zirconium is obtained using the two-step Van Arkel process. In the first step, impure zirconium and iodine are heated to produce zirconium iodide (ZrI_4). In the second step, ZrI_4 is decomposed to produce pure zirconium.

$$ZrI_4(s) \longrightarrow Zr(s) + 2I_2(g)$$

Determine the percent yield of zirconium if 45.0 g of ZrI_4 is decomposed and 5.00 g of pure Zr is obtained.

95. Methanol, wood alcohol, is produced when carbon monoxide reacts with hydrogen gas.

$$CO + 2H_2 \longrightarrow CH_3OH$$

When 8.50 g of carbon monoxide reacts with an excess of hydrogen, 8.52 g of methanol is collected. Complete **Table 11.4,** and calculate the percent yield.

Table 11.4 Methanol Reaction Data

	CO(g)	$CH_3OH(l)$
Mass	8.52 g	
Molar mass	28.01 g/mol	32.05 g/mol
Moles		

96. Phosphorus (P_4) is commercially prepared by heating a mixture of calcium phosphate ($CaSiO_3$), sand (SiO_2), and coke (C) in an electric furnace. The process involves two reactions.

$$2Ca_3(PO_4)_2(s) + 6SiO_2(s) \longrightarrow 6CaSiO_3(l) + P_4O_{10}(g)$$
$$P_4O_{10}(g) + 10C(s) \longrightarrow P_4(g) + 10CO(g)$$

The P_4O_{10} produced in the first reaction reacts with an excess of coke (C) in the second reaction. Determine the theoretical yield of P_4 if 250.0 g of $Ca_3(PO_4)_2$ and 400.0 g of SiO_2 are heated. If the actual yield of P_4 is 45.0 g, determine the percent yield of P_4.

97. Chlorine forms from the reaction of hydrochloric acid with manganese(IV) oxide. The balanced equation is:

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$$

Calculate the theoretical yield and the percent yield of chlorine if 86.0 g of MnO_2 and 50.0 g of HCl react. The actual yield of Cl_2 is 20.0 g.

Mixed Review

98. Ammonium sulfide reacts with copper(II) nitrate in a double replacement reaction. What mole ratio would you use to determine the moles of NH_4NO_3 produced if the moles of CuS are known?

99. Fertilizer The compound calcium cyanamide (CaNCN) is used as a nitrogen source for crops. To obtain this compound, calcium carbide is reacted with nitrogen at high temperatures.

$$CaC_2(s) + N_2(g) \longrightarrow CaNCN(s) + C(s)$$

What mass of CaNCN can be produced if 7.50 mol of CaC_2 reacts with 5.00 mol of N_2?

100. When copper(II) oxide is heated in the presence of hydrogen gas, elemental copper and water are produced. What mass of copper can be obtained if 32.0 g of copper(II) oxide is used?

101. Air Pollution Nitrogen oxide, which is present in urban air pollution, immediately converts to nitrogen dioxide as it reacts with oxygen.

a. Write the balanced chemical equation for the formation of nitrogen dioxide from nitrogen oxide.

b. What mole ratio would you use to convert from moles of nitrogen oxide to moles of nitrogen dioxide?

102. Electrolysis Determine the theoretical and percent yield of hydrogen gas if 36.0 g of water undergoes electrolysis to produce hydrogen and oxygen and 3.80 g of hydrogen is collected.

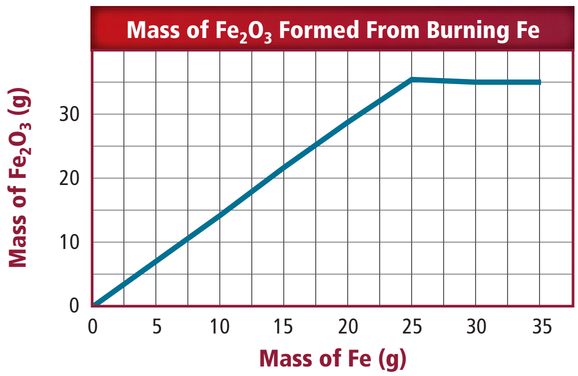

Figure 11.15

103. Iron reacts with oxygen as shown.

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

Different amounts of iron were burned in a fixed amount of oxygen. For each mass of iron burned, the mass of iron(II) oxide formed was plotted on the graph shown in **Figure 11.15.** Why does the graph level off after 25.0 g of iron is burned? How many moles of oxygen are present in the fixed amount?

Think Critically

104. Analyze and Conclude In an experiment, you obtain a percent yield of product of 108%. Is such a percent yield possible? Explain. Assuming that your calculation is correct, what reasons might explain such a result?

105. Observe and Infer Determine whether each reaction depends on a limiting reactant. Explain why or why not, and identify the limiting reactant.

a. Potassium chlorate decomposes to form potassium chloride and oxygen.

b. Silver nitrate and hydrochloric acid react to produce silver chloride and nitric acid.

106. Design an Experiment Design an experiment that can be used to determine the percent yield of anhydrous copper(II) sulfate when copper(II) sulfate pentahydrate is heated to remove water.

107. Apply When a campfire begins to die down and smolder, you can rekindle the flame by fanning the fire. Explain, in terms of stoichiometry, why the fire again begins to flare up when fanned.

108. Apply Students conducted a lab to investigate limiting and excess reactants. The students added different volumes of sodium phosphate solution (Na_3PO_4) to a beaker. They then added a constant volume of cobalt(II) nitrate solution ($Co(NO_3)_2$), stirred the contents, and allowed the beakers to sit overnight. The next day, each beaker had a purple precipitate at the bottom. The students decanted the supernatant from each beaker, divided it into two samples, and added one drop of sodium phosphate solution to one sample and one drop of cobalt(II) nitrate solution to the second sample. Their results are shown in **Table 11.5.**

a. Write a balanced chemical equation for the reaction.

b. Based on the results, identify the limiting reactant and the excess reactant for each trial.

Table 11.5 Reaction Data for $Co(NO_3)_2$ and Na_3PO_4

Trial	Volume Na_3PO_4	Volume $Co(NO_3)_2$	Reaction with Drop of Na_3PO_4	Reaction with Drop of $Co(NO_3)_2$
1	5.0 mL	10.0 mL	purple precipitate	no reaction
2	10.0 mL	10.0 mL	no reaction	purple precipitate
3	15.0 mL	10.0 mL	no reaction	purple precipitate
4	20.0 mL	10.0 mL	no reaction	purple precipitate

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Challenge Problem

109. When 9.59 g of a certain vanadium oxide is heated in the presence of hydrogen, water and a new oxide of vanadium are formed. This new vanadium oxide has a mass of 8.76 g. When the second vanadium oxide undergoes additional heating in the presence of hydrogen, 5.38 g of vanadium metal forms.
- **a.** Determine the empirical formulas for the two vanadium oxides.
- **b.** Write balanced equations for the steps of the reaction.
- **c.** Determine the mass of hydrogen needed to complete the steps of this reaction.

Cumulative Review

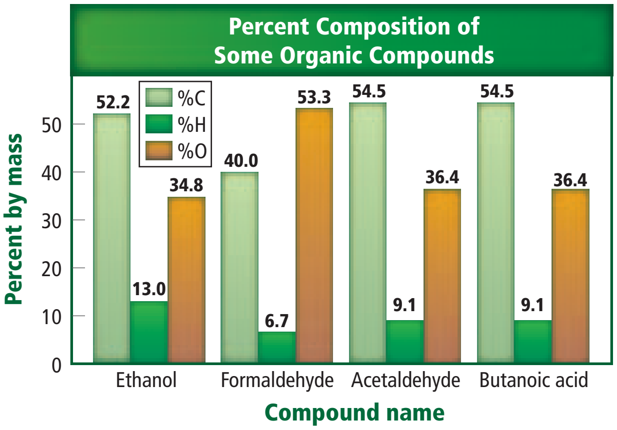
110. You observe that sugar dissolves more quickly in hot tea than in iced tea. You state that higher temperatures increase the rate at which sugar dissolves in water. Is this statement a hypothesis or a theory? Why? *(Chapter 1)*

111. Write the electron configuration for each of the following atoms. *(Chapter 5)*
- **a.** fluorine
- **b.** aluminum
- **c.** titanium
- **d.** radon

112. Explain why the gaseous nonmetals exist as diatomic molecules, but other gaseous elements exist as single atoms. *(Chapter 8)*

113. Write a balanced equation for the reaction of potassium with oxygen. *(Chapter 9)*

114. What is the molecular mass of UF_6? What is the molar mass of UF_6? *(Chapter 10)*



■ **Figure 11.16**

115. **Figure 11.16** gives percent composition data for several organic compounds. *(Chapter 10)*
- **a.** How are the molecular and empirical formulas of acetaldehyde and butanoic acid related?
- **b.** What is the empirical formula of butanoic acid?

Additional Assessment

WRITING in Chemistry

116. Air Pollution Research the air pollutants produced by combustion of gasoline in internal combustion engines. Discuss the common pollutants and the reaction that produces them. Show, through the use of stoichiometry, how each pollutant could be reduced if more people used mass transit.

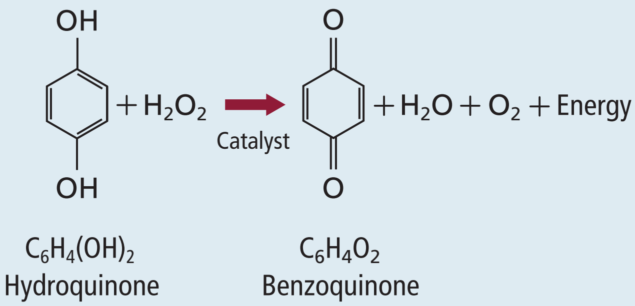
117. Haber Process The percent yield of ammonia produced when hydrogen and nitrogen are combined under ordinary conditions is extremely small. However, the Haber Process combines the two gases under a set of conditions designed to maximize yield. Research the conditions used in the Haber Process, and find out why the development of the process was of great importance.

DBQ Document-Based Question

Chemical Defense *Many insects secrete hydrogen peroxide (H_2O_2) and hydroquinone $C_6H_4(OH)_2$. Bombardier beetles take this a step further by mixing these chemicals with a catalyst. The result is an exothermic chemical reaction and a spray of hot, irritating chemicals for any would-be predator. Researchers hope to use a similar method to reignite aircraft turbine engines.*

Figure 11.17 *below shows the unbalanced chemical reaction that results in the bombardier beetle's defensive spray.*

Data obtained from: Becker, Bob. April 2006. *ChemMatters*. 24: no. 2.



■ **Figure 11.17**

118. Balance the equation in **Figure 11.17.** If the bombardier beetle stores 100.0 mg of hydroquinone ($C_6H_4(OH)_2$) along with 50.0 mg of hydrogen peroxide (H_2O_2), what is the limiting reactant?

119. What is the excess reactant and how many milligrams are in excess?

120. How many milligrams of benzoquinone will be produced?

Cumulative

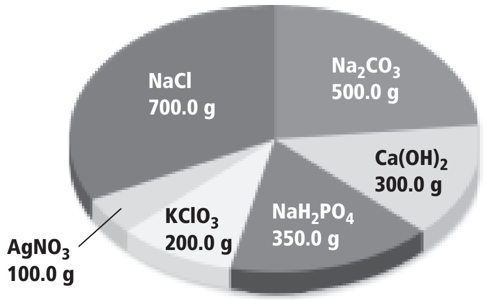
Standardized Test Practice

Multiple Choice

1. Stoichiometry is based on the law of
- **A.** constant mole ratios.
- **B.** Avogadro's constant.
- **C.** conservation of energy.
- **D.** conservation of mass.

Use the graph below to answer Questions 2 to 5.

Supply of Various Chemicals in Dr. Raitano's Laboratory



2. Pure silver metal can be made using the reaction shown below.

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

How many grams of copper metal will be needed to use up all of the $AgNO_3$ in Dr. Raitano's laboratory?
- **A.** 18.70 g
- **B.** 37.3 g
- **C.** 74.7 g
- **D.** 100 g

3. The LeBlanc process is the traditional method of manufacturing sodium hydroxide. The equation for this process is as follows.

$$Na_2CO_3(aq) + Ca(OH)_2(aq) \rightarrow 2NaOH(aq) + CaCO_3(s)$$

Using the amounts of chemicals available in Dr. Raitano's lab, what is the maximum number of moles of NaOH that can be produced?
- **A.** 4.05 mol
- **B.** 4.72 mol
- **C.** 8.097 mol
- **D.** 9.43 mol

4. Pure O_2 gas can be generated from the decomposition of potassium chlorate ($KClO_3$):

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

If half of the $KClO_3$ in the lab is used and 12.8 g of oxygen gas is produced, what is the percent yield of this reaction?
- **A.** 12.8%
- **B.** 32.7%
- **C.** 65.6%
- **D.** 98.0%

5. Sodium dihydrogen pyrophosphate ($Na_2H_2P_2O_7$), more commonly known as baking powder, is manufactured by heating NaH_2PO_4 to a high temperature.

$$2NaH_2PO_4(s) \rightarrow Na_2H_2P_2O_7(s) + H_2O(g)$$

If 444.0 g of $Na_2H_2P_2O_7$ is needed, how much more NaH_2PO_4 will Dr. Raitano have to buy to make enough $Na_2H_2P_2O_7$?
- **A.** 0.00 g
- **B.** 94.0 g
- **C.** 130.0 g
- **D.** 480 g

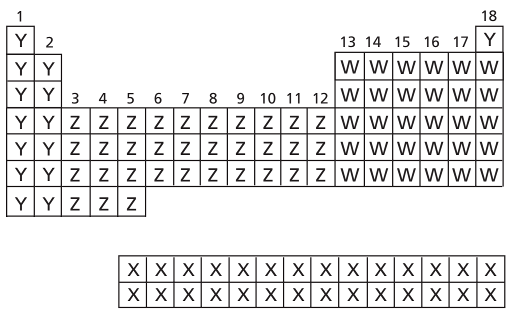
6. Red mercury(II) oxide decomposes at high temperatures to form mercury metal and oxygen gas.

$$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$$

If 3.55 mol of HgO decomposes to form 1.54 mol of O_2 and 618 g of Hg, what is the percent yield of this reaction?
- **A.** 13.2%
- **B.** 42.5%
- **C.** 56.6%
- **D.** 86.8%

Use the diagram below to answer Questions 7 and 8.

PERIODIC TABLE



7. Which elements tend to have the largest atomic radius in their periods?
- **A.** W
- **B.** X
- **C.** Y
- **D.** Z

8. Elements labeled *W* have their valence electrons in which sublevel?
- **A.** s
- **B.** p
- **C.** d
- **D.** f

Short Answer

9. Dimethyl hydrazine $(CH_3)_2N_2H_2$ ignites on contact with dinitrogen tetroxide (N_2O_4).

$$(CH_3)_2N_2H_2(l) + 2N_2O_4(l) \longrightarrow 3N_2(g) + 4H_2O(g) + 2CO_2(g)$$

Because this reaction produces an enormous amount of energy from a small amount of reactants, it was used to drive the rockets on the Lunar Excursion Modules (LEMs) of the Apollo space program. If 18.0 mol of dinitrogen tetroxide is consumed in this reaction, how many moles of nitrogen gas will be released?

Extended Response

Use the table below to answer Questions 10 and 11.

First Ionization Energy of Period 3 Elements		
Element	Atomic Number	1st Ionization Energy, kJ/mol
Sodium	11	496
Magnesium	12	736
Aluminum	13	578
Silicon	14	787
Phosphorus	15	1012
Selenium	16	1000
Chlorine	17	1251
Argon	18	1521

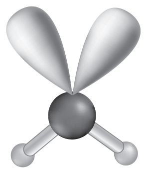
10. Plot the data from this data table. Place atomic numbers on the *x*-axis.

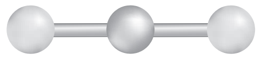
11. Summarize the general trend in ionization energy. How does ionization energy relate to the number of valence electrons in an element?

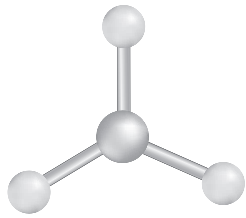
SAT Subject Test: Chemistry

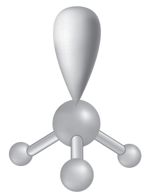
12. How much cobalt(III) titanate (CO_2TiO_4), in moles, is in 7.13 g of the compound?
A. 2.39×10^1 mol
B. 3.10×10^{-2} mol
C. 3.22×10^1 mol
D. 4.17×10^{-2} mol
E. 2.28×10^{-2} mol

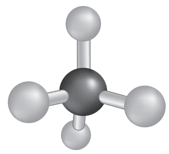
Use the pictures below to answer Questions 13 to 17.

A.


D.


B.


E.


C.


13. Hydrogen sulfide displays this molecular shape.

14. Molecules with this shape have four shared pairs of electrons and no lone pairs of electrons.

15. This molecular shape is known as *trigonal planar*.

16. Carbon dioxide displays this molecular shape.

17. This molecular shape undergoes sp^2 hybridization.

NEED EXTRA HELP?																	
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
Review Section . . .	11.1	11.2	11.2	11.4	11.3	11.4	6.3	5.3	11.2	6.3	6.3	10.3	8.4	8.4	8.4	8.4	8.4