Stoichiometry

What You'll Learn

- You will write mole ratios from balanced chemical equations.
- You will calculate the number of moles and the mass of a reactant or product when given the number of moles or the mass of another reactant or product.
- You will identify the limiting reactant in a chemical reaction.
- You will determine the percent yield of a chemical reaction.

Why It's Important

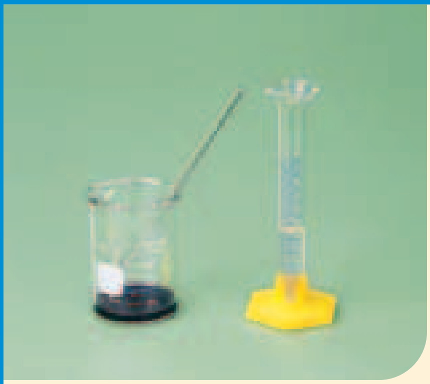
The cost of the things you buy is lower because chemists use stoichiometric calculations to increase efficiency in laboratories, decrease waste in manufacturing, and produce products more quickly.

Chemistry Online

Visit the Chemistry Web site at **chemistrymc.com** to find links to stoichiometry.

The candle will continue to burn as long as oxygen and candle wax are present.

DISCOVERY LAB



Observing a Chemical Reaction

Reactants are consumed in a chemical reaction as products are produced. What evidence can you observe that a reaction takes place?

Safety Precautions

Materials

10-mL graduated cylinder
100-mL beaker
stirring rod
0.01*M* potassium permanganate ($KMnO_4$)
0.01*M* sodium hydrogen sulfite ($NaHSO_3$)

Procedure

1. Measure 5.0 mL of 0.01*M* potassium permanganate solution ($KMnO_4$) and pour it into a 100-mL beaker.
2. Add 5.0 mL of 0.01*M* sodium hydrogen sulfite solution ($NaHSO_3$) to the potassium permanganate solution while stirring. Record your observations.
3. Slowly add additional 5.0-mL portions of the $NaHSO_3$ solution until the $KMnO_4$ solution turns colorless. Record your observations.
4. Record the total volume of the $NaHSO_3$ solution you used to cause the beaker's contents to become colorless.

Analysis

What evidence do you have that a reaction occurred? Would anything more have happened if you continued to add $NaHSO_3$ solution to the beaker? Explain.

Section 12.1 What is stoichiometry?

Objectives

- **Identify** the quantitative relationships in a balanced chemical equation.
- **Determine** the mole ratios from a balanced chemical equation.

Vocabulary

stoichiometry
mole ratio

Were you surprised when, in doing the **DISCOVERY LAB,** you saw the purple color of potassium permanganate disappear 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. The photo on the opposite page shows the combustion of a candle using the oxygen in the surrounding air. What would happen if a bell jar was lowered over the burning candle blocking off the supply of oxygen? You know that oxygen is needed for the combustion of candle wax, so when the oxygen inside the bell jar is used up, the candle will go out.

Chemical reactions, such as the reaction of potassium permanganate with sodium hydrogen sulfite and the combustion of a candle, stop when one of the reactants is used up. Thus, in planning the reaction of potassium permanganate and sodium hydrogen sulfite, a chemist needs to know how many grams of potassium permanganate are needed to react completely with a known mass of sodium hydrogen sulfite. You might ask, "How much oxygen is required to completely burn a candle of known mass, or how much product will be produced if a given amount of a reactant is used?" Stoichiometry is the tool for answering these questions.

Mole-Mass Relationships in Chemical Reactions

The study of quantitative relationships between amounts of reactants used and products formed by a chemical reaction is called **stoichiometry**. Stoichiometry is based on the law of conservation of mass, which was introduced by Antoine Lavoisier in the eighteenth century. The law states that matter is neither created nor destroyed in a chemical reaction. Chemical bonds in reactants break and new chemical bonds form to produce products, but the amount of matter present at the end of the reaction is the same as was present at the beginning. Therefore, the mass of the reactants equals the mass of the products.

Careers Using Chemistry

Pharmacist

Are you interested in taking an active role in the health care of others? Would you like to advise physicians as well as patients? Then consider a career as a pharmacist.

Pharmacists must understand the composition and use of prescribed drugs and medicines, and over-the-counter medications. They advise doctors and patients about proper use, harmful combinations, and possible side-effects. Although pharmaceutical companies supply most medicines, pharmacists may do the actual mixing of ingredients to form powders, tablets, capsules, ointments, and solutions.

Stoichiometry and the balanced chemical equation Look at the reaction of powdered iron with oxygen shown in **Figure 12-1**. As tiny particles of iron react with oxygen in the air, iron(III) oxide (Fe_2O_3) is produced.

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

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. But, 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.

Does the chemical equation tell you anything about the masses of the reactants and products? Not directly. But as you learned in Chapter 11, the mass of any substance can be determined by multiplying the number of moles of the substance by the conversion factor that relates mass and number of moles, which is the molar mass. Thus, the mass of the reactants can be calculated in this way.

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

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

The total mass of the reactants = 319.4 g

Similarly, the mass of the product is

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

The total mass of the reactants equals the mass of the product, as predicted by the law of conservation of mass. **Table 12-1** summarizes the relationships that can be determined from a balanced chemical equation.

Figure 12-1

If you know the equation for this reaction between iron and oxygen, you can calculate the number of moles and the mass of each reactant and product.

Table 12-1

Relationships Derived from a Balanced Chemical Equation				
Iron	+	**Oxygen**	→	**Iron(III) oxide**
4Fe(s)	+	$3O_2(g)$	→	$2Fe_2O_3(s)$
4 atoms Fe	+	3 molecules O_2	→	2 formula units Fe_2O_3
4 moles Fe	+	3 moles O_2	→	2 moles Fe_2O_3
223.4 g Fe	+	96.0 g O_2	→	319.4 g Fe_2O_3
		319.4 g reactants	→	319.4 g product

EXAMPLE PROBLEM 12-1

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 formulas in the equation represent both representative particles (molecules) and moles. 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

The equation in terms of molecules = ?
The equation in terms of moles = ?
The equation in terms of mass = ?

2. Solve for the Unknown

The coefficients 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 indicate the number of moles.

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

Calculate the mass of each reactant and product by multiplying the number of moles by the conversion factor molar mass.

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

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

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

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

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

Add the masses of the reactants.

44.09 g C_3H_8 + 160.0 g O_2 = 204.1 g reactants

Add the masses of the products.

132.0 g CO_2 + 72.08 g H_2O = 204.1 g 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.

Because propane gas is readily liquified, it can be stored in tanks and transported to wherever it is needed.

Practice!

For more practice interpreting chemical equations, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

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

b. $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$

c. $4Zn(s) + 10HNO_3(aq) \rightarrow 4Zn(NO_3)_2(aq) + N_2O(g) + 5H_2O(l)$

d. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

e. $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Mole ratios You have seen that the coefficients in a chemical equation indicate the relationships among moles of reactants and products. For example, return to the reaction between iron and oxygen described in **Table 12-1**. The equation indicates that four moles of iron react with three moles of oxygen. It also indicates that four moles of iron react to produce two moles of iron(III) oxide. How many moles of oxygen react to produce two moles of iron(III) oxide? You can use the relationships between coefficients to write conversion factors called mole ratios. A **mole ratio** is a ratio between the numbers of moles of any two substances in a balanced chemical equation. As another example, consider the reaction shown in **Figure 12-2**. Aluminum reacts with bromine to form aluminum bromide. Aluminum bromide is used as a catalyst to speed up a variety of chemical reactions.

$$2Al(s) + 3Br_2(l) \rightarrow 2AlBr_3(s)$$

Figure 12-2

Bromine is one of the two elements that are liquids at room temperature. Mercury is the other. Aluminum is a lightweight metal that resists corrosion. Aluminum and bromine react vigorously to form the ionic compound aluminum bromide.

What mole ratios can be written for this reaction? Starting with the reactant aluminum, you can write a mole ratio that relates the moles of aluminum to the moles of bromine. Another mole ratio shows how the moles of aluminum relate to the moles of aluminum bromide.

$$\frac{2 \text{ mol Al}}{3 \text{ mol Br}_2} \quad \text{and} \quad \frac{2 \text{ mol Al}}{2 \text{ mol AlBr}_3}$$

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

$$\frac{3 \text{ mol Br}_2}{2 \text{ mol Al}} \quad \text{and} \quad \frac{3 \text{ mol Br}_2}{2 \text{ mol AlBr}_3}$$

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

$$\frac{2 \text{ mol AlBr}_3}{2 \text{ mol Al}} \quad \text{and} \quad \frac{2 \text{ mol AlBr}_3}{3 \text{ mol Br}_2}$$

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.

What mole ratios can be written for the decomposition of potassium chlorate ($KClO_3$)? This reaction is sometimes used to obtain small amounts of oxygen in the laboratory.

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

Each substance forms a mole ratio with the two other substances in the reaction. Thus, each substance should be the numerator of two mole ratios.

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

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

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

Can you explain why each reactant and product is in the numerator two times?

Six mole ratios define all the relationships in this reaction, which has three participating species. How many could you write for a reaction involving a total of four reactants and products? A simple way to find out is to multiply the number of species in the equation by the next lower number. Thus, you could write 12 mole ratios for a reaction involving four species.

PRACTICE PROBLEMS

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

3. 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

Practice!

For more practice writing mole ratios, go to **Supplemental Practice Problems** in Appendix A.

You may be wondering why you need to learn to write mole ratios. As you will see in the next section, mole ratios are the key to calculations based upon a chemical equation. Suppose you know the amount of one reactant you will use in a chemical reaction. With the chemical equation and the mole ratios, you can calculate the amount of any other reactant in the equation and the maximum amount of product you can obtain.

Section 12.1 Assessment

4. What is stoichiometry?

5. List three ways in which a balanced chemical equation can be interpreted.

6. What is a mole ratio?

7. **Thinking Critically** Write a balanced chemical equation for each reaction and determine the possible mole ratios.

a. Nitrogen reacts with hydrogen to produce ammonia.

b. Hydrogen peroxide (H_2O_2) decomposes to produce water and oxygen.

c. Pieces of zinc react with a phosphoric acid solution to produce solid zinc phosphate and hydrogen gas.

8. **Formulating Models** Use the balanced chemical equation to determine the mole ratios for the reaction of hydrogen and oxygen, $2H_2(g) + O_2(g) \longrightarrow 2H_2O$. Make a drawing showing six molecules of hydrogen reacting with the correct number of oxygen molecules. Show the number of molecules of water produced.

Section 12.2

Stoichiometric Calculations

Objectives

- **Explain** the sequence of steps used in solving stoichiometric problems.
- **Use** the steps to solve stoichiometric problems.

Suppose a chemist needs to obtain a certain amount of product from a reaction. How much reactant must be used? Or, suppose the chemist wants to know how much product will form if a certain amount of reactant is used. Chemists use stoichiometric calculations to answer these questions.

Using Stoichiometry

Recall that stoichiometry is the study of quantitative relationships between the amounts of reactants used and the amounts of products formed by a chemical reaction. What are the tools needed for stoichiometric calculations? All stoichiometric calculations begin with a balanced chemical equation, which indicates relative amounts of the substances that react and the products that form. Mole ratios based on the balanced chemical equation are also needed. You learned to write mole ratios in Section 12.1. Finally, mass-to-mole conversions similar to those you learned about in Chapter 11 are required.

Stoichiometric mole-to-mole conversion The vigorous reaction between potassium and water is shown in **Figure 12-3**. How can you determine the number of moles of hydrogen produced when 0.0400 mole of potassium is used? Start by writing the balanced chemical equation.

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

Then, identify the substance that you know and the substance that you need to determine. The given substance is 0.0400 mole of potassium. The unknown is the number of moles of hydrogen. Because the quantity of the given substance is in moles and the unknown substance is to be determined in moles, this problem is 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 12.1 you learned to use the balanced chemical equation to write mole ratios that describe mole relationships. Mole ratios are used as conversion factors to convert a known number of moles of one substance to moles of another substance in the same chemical reaction. What mole ratio could be used to convert moles of potassium to moles of hydrogen? In the correct mole ratio, the moles of unknown (H_2) should be the numerator and the moles of known (K) should be the denominator. The correct mole ratio is

$$\frac{1 \text{ mol } H_2}{2 \text{ mol K}}$$

This mole ratio can be used to convert the known number of moles of potassium to a number of moles of hydrogen. Remember that when you use a conversion factor, the units must cancel.

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

$$0.0400 \cancel{\text{mol K}} \times \frac{1 \text{ mol } H_2}{2 \cancel{\text{mol K}}} = 0.0200 \text{ mol } H_2$$

If you put 0.0400 mol K into water, 0.0200 mol H_2 will be produced. The **How It Works** feature at the end of this chapter shows the importance of mole ratios.

Figure 12–3

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

EXAMPLE PROBLEM 12-2

Stoichiometric Mole-to-Mole Conversion

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 growing concentration of CO_2 in the atmosphere. How many moles of carbon dioxide are produced when 10.0 moles of propane are burned in excess oxygen in a gas grill?

1. Analyze the Problem

You are given moles of the reactant propane, and moles of the product carbon dioxide must be found. The balanced chemical equation must be written. Conversion from moles of C_3H_8 to moles of CO_2 is required. The correct mole ratio has moles of unknown substance in the numerator and moles of known substance in the denominator.

Known

moles of propane = 10.0 mol C_3H_8

Unknown

moles of carbon dioxide = ? mol CO_2

2. Solve for the Unknown

Write the balanced chemical equation. Label the known substance and the unknown substance.

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

Determine the mole ratio that relates mol CO_2 to mol C_3H_8.

$$\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$$

Multiply the known number of moles of C_3H_8 by the mole ratio.

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

Burning 10.0 mol C_3H_8 produces 30.0 mol CO_2.

3. Evaluate the Answer

The given number of moles has three significant figures. Therefore, the answer must have three digits. The balanced chemical equation indicates that 1 mol C_3H_8 produces 3 mol CO_2. Thus, 10.0 mol C_3H_8 would produce three times as many moles of CO_2, or 30.0 mol.

Math Handbook

Review dimensional analysis in the **Math Handbook** on page 900 of this text.

PRACTICE PROBLEMS

9. Sulfuric acid is formed when sulfur dioxide reacts with oxygen and water. Write the balanced chemical equation for the reaction. If 12.5 mol SO_2 reacts, how many mol H_2SO_4 can be produced? How many mol O_2 is needed?

10. A reaction between methane and sulfur produces 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 mol CS_2 produced when 1.50 mol S_8 is used.

c. How many mol H_2S is produced?

Practice!

For more practice converting from moles of one substance to moles of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

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 situation is an example of a mole-to-mass conversion.

The reaction of sodium and chlorine to form sodium chloride releases a large amount of energy in the form of light and heat. It should not surprise you, then, that a large amount of energy is required to decompose sodium chloride.

EXAMPLE PROBLEM 12-3

Stoichiometric Mole-to-Mass Conversion

Determine the mass of sodium chloride or table salt (NaCl) produced when 1.25 moles of chlorine gas reacts vigorously with 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

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

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

Write the mole ratio that relates mol NaCl to mol Cl_2.

$$\frac{2\ \text{mol NaCl}}{1\ \text{mol } Cl_2}$$

Multiply the number of moles of Cl_2 by the mole ratio.

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

Multiply mol NaCl by the molar mass of NaCl.

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

3. Evaluate the Answer

The given number of moles has three significant figures, so the mass of NaCl is correctly stated with three digits. The computations are correct and the unit is as expected.

Practice!

For more practice converting from moles of one substance to mass of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

11. 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 using chlorine and coke (carbon).

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

If you begin with 1.25 mol TiO_2, what mass of Cl_2 gas is needed?

12. Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How many grams of chlorine gas can be obtained from 2.50 mol NaCl?

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 12-4 will demonstrate 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 determining a mole ratio.

Try at Home LAB

See page 957 in Appendix E for **Baking Soda Stoichiometry**

EXAMPLE PROBLEM 12-4

Stoichiometric Mass-to-Mass Conversion

Ammonium nitrate (NH_4NO_3), an important fertilizer, produces N_2O gas and H_2O when it decomposes. Determine the mass of water produced from the decomposition of 25.0 g of solid ammonium nitrate.

The vigorous decomposition of ammonium nitrate attests to its use in explosives. However, ammonium nitrate is also widely used as a fertilizer because it is 100 percent available for plant use.

1. Analyze the Problem

You are given the mass of the reactant and will need to write the balanced chemical equation. You then must convert from the mass of the reactant to moles of the reactant. You will next use a mole ratio to relate moles of the reactant to moles of the product. Finally, you will 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

Write the balanced chemical equation for the reaction and identify the known and unknown substances.

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

Convert grams of NH_4NO_3 to moles of NH_4NO_3 using the inverse of molar mass as the conversion factor.

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

Determine from the equation the mole ratio of mol H_2O to mol NH_4NO_3. The unknown quantity is the numerator.

$$\frac{2 \text{ mol } H_2O}{1 \text{ mol } NH_4NO_3}$$

Multiply mol NH_4NO_3 by the mole ratio.

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

Calculate the mass of H_2O using molar mass as the conversion factor.

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

3. Evaluate the Answer

The number of significant figures in the answer, three, is determined by the given moles of ammonium nitrate. The calculations are correct and the unit is appropriate.

Practice!

For more practice converting from mass of one substance to mass of another substance in a chemical equation, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

13. One in a series of reactions that inflate air bags in automobiles is the decomposition of sodium azide (NaN_3).

$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$

Determine the mass of N_2 produced if 100.0 g NaN_3 is decomposed.

14. In the formation of acid rain, sulfur dioxide reacts with oxygen and water in the air to form sulfuric acid. Write the balanced chemical equation for the reaction. If 2.50 g SO_2 react with excess oxygen and water, how many grams of H_2SO_4 are produced?

The steps you followed in Example Problem 12-4 are illustrated in **Figure 12-4** and described below it. Use the steps as a guide when you do stoichiometric calculations until you become thoroughly familiar with the procedure. Study **Figure 12-4** as you read.

The specified unit of the given substance determines at what point you will start your calculations. If the amount of the given substance is in moles, step 2 is omitted and step 3, mole-to-mole conversion, becomes the starting point for the calculations. However, if mass is the starting unit, calculations begin with step 2. The end point of the calculation depends upon the specified unit of the unknown substance. If the answer is to be obtained in moles, the calculation is finished with step 3. If the mass of the unknown is to be determined, you must go on to step 4.

Like any other type of problem, stoichiometric calculations require practice. You can begin to practice your skills in the **miniLAB** that follows.

miniLAB

Baking Soda Stoichiometry

Predicting When baking soda is an ingredient in your recipe, its purpose is to make the batter rise and produce a product with a light and fluffy texture. That's because baking soda, or sodium hydrogen carbonate ($NaHCO_3$), decomposes upon heating to form carbon dioxide gas.

$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

Predict how much sodium carbonate (Na_2CO_3) is produced when baking soda decomposes.

Materials ring stand, ring, clay triangle, crucible, crucible tongs, Bunsen burner, balance, 3.0 g baking soda ($NaHCO_3$)

Procedure

1. Measure the mass of a clean, dry crucible. Add about 3.0 g of $NaHCO_3$ and measure the combined mass of the crucible and $NaHCO_3$. Record both masses in your data table and calculate the mass of the $NaHCO_3$.
2. Use this starting mass of baking soda and the balanced chemical equation to calculate the mass of Na_2CO_3 that will be produced.
3. Set up a ring stand with a ring and clay triangle for heating the crucible.
4. Heat the crucible slowly at first and then with a stronger flame for 7–8 min. Use tongs to remove the hot crucible. Record your observations during the heating.
5. Allow the crucible to cool and then obtain the mass of the crucible and sodium carbonate.

Analysis

1. What were your observations during the heating of the baking soda?
2. How did your calculated mass of sodium carbonate compare with the actual mass you obtained from the experiment? If the two masses are different, suggest reasons for the difference.

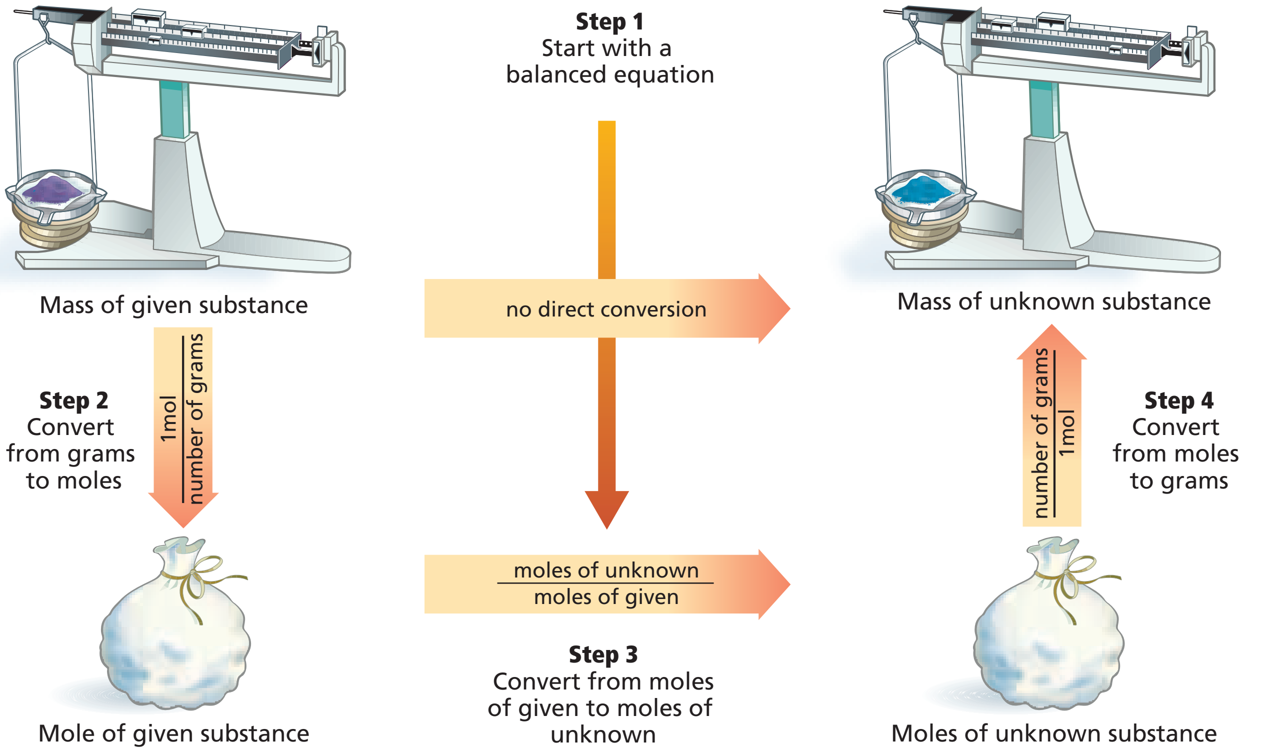


Figure 12-4

Follow the steps from the balanced equation to the mass of the unknown. Note that there is no shortcut from the mass of the given substance to the mass of the unknown substance. The route goes through the mole. However, you can follow the arrow from step 1 to step 3 if the amount of the given substance is in moles.

Steps in Stoichiometric Calculations

1. *Write a balanced chemical equation.* Interpret the equation in terms of moles.
2. *Determine the moles of the given substance using a mass-to-mole conversion.* Use the inverse of the molar mass as the conversion factor.
3. *Determine the moles of the unknown substance from the moles of the given substance.* Use the appropriate mole ratio from the balanced chemical equation as the conversion factor.
4. *From the moles of the unknown substance, determine the mass of the unknown substance using a mole-to-mass conversion.* Use the molar mass as the conversion factor.

Section 12.2 Assessment

15. Why is a balanced chemical equation needed in solving stoichiometric calculations?

16. When solving stoichiometric problems, how is the correct mole ratio expressed?

17. List the four steps used in solving stoichiometric problems.

18. **Thinking Critically** In a certain industrial process, magnesium reacts with liquid bromine. How would a chemical engineer determine the mass of bromine needed to react completely with a given mass of magnesium?

19. **Concept Mapping** Many cities use calcium chloride to prevent ice from forming on roadways. To produce calcium chloride, calcium carbonate (limestone) is reacted with hydrochloric acid according to this equation.

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

Create a concept map that describes how you can determine the mass of calcium chloride produced if the mass of hydrochloric acid is given.

Section 12.3 Limiting Reactants

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.

Vocabulary

limiting reactant
excess reactant

At a school dance, the music begins and boys and girls pair up to dance. If there are more boys than girls, some boys will be left without partners. The same is true of reactants in a chemical reaction. Rarely in nature are reactants in a chemical reaction present in the exact ratios specified by the balanced equation. Generally, one or more reactants are in excess and the reaction proceeds until all of one reactant is used up.

Why do reactions stop?

When a chemical reaction is carried out in the laboratory, the same principle applies. Usually, one or more reactants are in excess, while one is limited. The amount of product depends upon the reactant that is limited.

Remember the reaction between potassium permanganate and sodium hydrogen sulfite in the **DISCOVERY LAB.** As you added colorless sodium hydrogen sulfite to purple potassium permanganate, the color faded as a reaction took place. Finally, the solution was colorless. You could have continued adding sodium hydrogen sulfite, but would any further reaction have taken place? You are correct if you said that no further reaction could take place because no potassium permanganate was available to react. 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. A portion of all of the other reactants remains after the reaction stops. These left-over reactants are called **excess reactants**. What was the excess reactant in the reaction of potassium permanganate and sodium hydrogen sulfite?

To help you understand limiting reactants, consider the analogy in **Figure 12-5**. How many tool sets can be assembled from the items shown if each complete tool set consists of one pair of pliers, one hammer, and two screwdrivers? You can see that four complete tool sets can be assembled. The number of tool sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess. Chemical reactions work in a similar way.

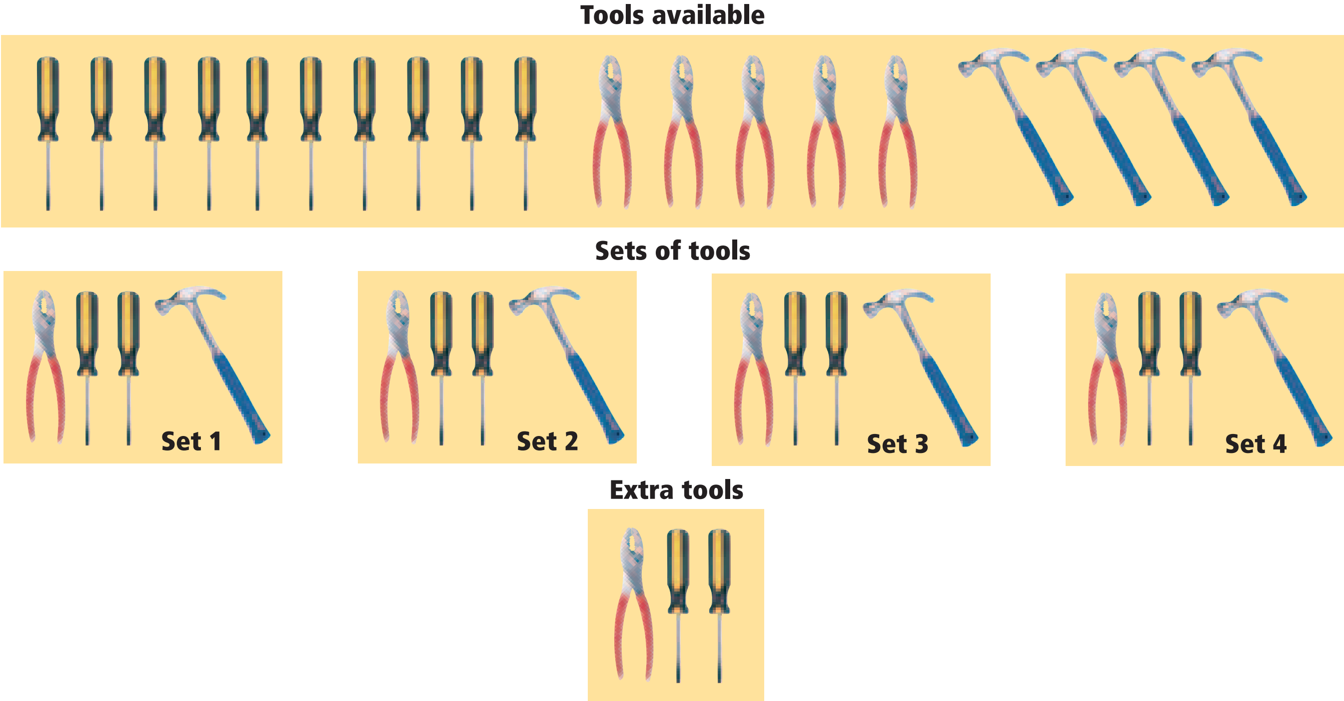
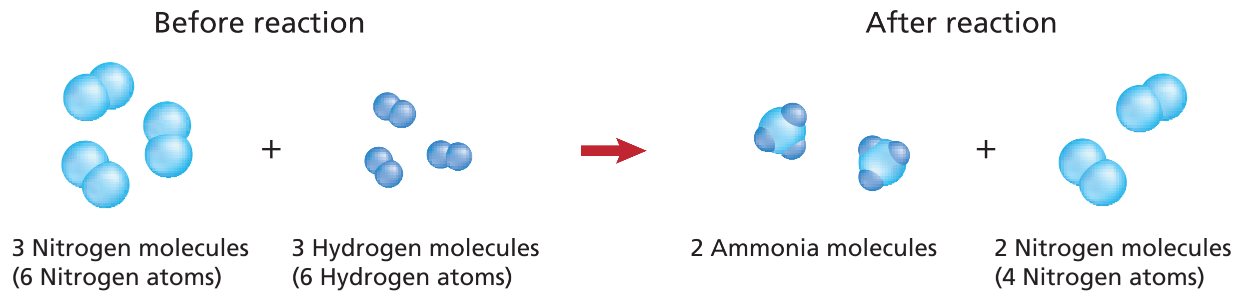


Figure 12-5

Each tool set must have one hammer so only four sets can be assembled. Which tool is limiting? Which tools are in excess?



The calculations you did in Section 12.2 were based on having the reactants present in the ratio described by the balanced chemical equation. How can you calculate the amount of product formed when one reactant limits the amount of product and the other is in excess? The first thing you must do is determine which reactant is the limiting reactant.

Consider the reaction shown in **Figure 12-6** in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). You can visualize that in the first step of the reaction, all the nitrogen molecules and hydrogen molecules are separated into individual atoms. These are the atoms available for reassembling into ammonia molecules just like the tools in **Figure 12-5** before they were assembled into tool kits. How many molecules of ammonia will be produced from the available atoms? Four tool kits could be assembled from the tools because only four hammers were available. Two ammonia molecules can be assembled from the hydrogen and nitrogen atoms because only six hydrogen atoms are available, three for each ammonia molecule. When the hydrogen is gone, two molecules of nitrogen remain unreacted. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It's important to know which reactant is the limiting reactant because, as you have just learned, the amount of product formed depends upon this reactant.

Figure 12-6

Refer to **Table C-1** in Appendix C for a key to atom color conventions. Check to see whether all the atoms present before the reaction are present after the reaction. Some nitrogen molecules were unchanged in the reaction. Which reactant is in excess?

Calculating the Product When a Reactant Is Limited

How can you determine which reactant is limited? As an example, consider the formation of disulfur dichloride (S_2Cl_2). Disulfur dichloride is used to vulcanize rubber, a process that makes rubber harder, stronger, and less likely to become soft when hot or brittle when cold. In the production of disulfur dichloride, molten sulfur reacts with chlorine gas according to this 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?

Masses of both reactants are given. You must first determine which one is the limiting reactant because the reaction will stop producing product when the limiting reactant is used up. Identifying the limiting reactant involves finding the number of moles of each reactant. This is done by converting the masses of chlorine and sulfur to moles. Multiply each mass by the conversion factor that relates moles and mass, the inverse of the molar mass.

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

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

Chemistry Online

Topic: Limiting Reactants
To learn more about limiting reactants, visit the Chemistry Web site at **chemistrymc.com**

Activity: Research and explain how nutrients in soil can be depleted when crops are grown on a plot of land year after year. What is the limiting reactant in this case? How may the soil be enriched?

Biology CONNECTION

Vitamins and minerals are needed in small amounts in your daily diet. These substances are essential to the reactions in normal metabolism, growth and development, and regulation of the functioning of your body's cells.

Vitamin B-12 is necessary for red blood cells to mature properly in the bone marrow. People with pernicious anemia lose their ability to make intrinsic factor, a substance secreted by the stomach's membrane lining that enables vitamin B-12 to be absorbed from the intestine. Without an adequate amount of vitamin B-12, the body is unable to synthesize DNA properly, which affects red blood cell production.

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 four moles of chlorine are needed to react with one mole 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 available moles of chlorine by the available moles of sulfur.

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

Only 1.808 mol of chlorine is actually 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.

After the limiting reactant has been determined, 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 that relates disulfur dichloride and chlorine.

$$1.410 \cancel{\text{mol } Cl_2} \times \frac{4 \text{ mol } S_2Cl_2}{4 \cancel{\text{mol } Cl_2}} = 1.410 \text{ mol } S_2Cl_2$$

Then, moles of S_2Cl_2 is converted to grams of S_2Cl_2 by multiplying by the conversion factor that relates mass and moles, molar mass.

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

These two calculations can be combined into one like this.

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

Now you know that 190.4 g S_2Cl_2 is produced when 1.410 mol Cl_2 reacts with an excess of S_8.

What about the reactant sulfur, which you know is in excess? How much of it actually reacted? You can calculate the mass of sulfur needed to react completely with 1.410 mol of chlorine using a mole-to-mass calculation. The first step is to multiply the moles of chlorine by the mole ratio of sulfur to chlorine to obtain the number of moles of sulfur. Remember, the unknown is the numerator and the known is the denominator.

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

Now, to obtain the mass of sulfur needed, 0.3525 mol S_8 is multiplied by the conversion factor that relates mass and moles, molar mass.

$$0.3525 \cancel{\text{mol } S_8} \times \frac{256.5 \text{ g } S_8}{1 \cancel{\text{mol } S_8}} = 90.42 \text{ g } S_8 \text{ needed}$$

Knowing that 90.42 g S_8 is needed, you can calculate the amount of sulfur left unreacted when the reaction ends. Because 200.0 g of sulfur is available and only 90.42 g is needed, the mass in excess is

$$200.0 \text{ g } S_8 \text{ available} - 90.42 \text{ g } S_8 \text{ needed} = 109.6 \text{ g } S_8 \text{ in excess.}$$

EXAMPLE PROBLEM 12-5

Determining the Limiting Reactant

The reaction between solid white phosphorus 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 tetraphosphorus decoxide formed if 25.0 g of phosphorus (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 the limiting reactant must be identified and used for finding the mass of product. From moles of limiting reactant, the moles of the excess reactant used in the reaction can be determined. The number of moles of excess reactant that actually reacted can be converted to mass and subtracted from the given mass to find the amount in excess.

Known

mass of phosphorus = 25.0 g P_4

mass of oxygen = 50.0 g O_2

Unknown

mass of tetraphosphorus decoxide = ? g P_4O_{10}

mass of excess reactant = ? g excess reactant

2. Solve for the Unknown

a. Write the balanced chemical equation and identify the knowns and the unknown.

$$\begin{array}{ccccc} 25.0\text{ g} & & 50.0\text{ g} & \rightarrow & ?\text{ g} \\ P_4(s) & + & 5O_2(g) & \rightarrow & P_4O_{10}(s) \end{array}$$

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.

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

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

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

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

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

$$\frac{5\text{ mol } O_2}{1\text{ mol } P_4}$$

Because 7.72 mol O_2 is available but only 5 mol is needed to react with 1 mol 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).

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

Continued on next page

Tiny pieces of white phosphorus deposited on filter paper burst into flame on contact with air. That's why white phosphorus is never found free in nature. Phosphorus is an essential element in living systems; for example, phosphate groups occur regularly along strands of DNA.

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.

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

b. 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.

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

Multiply moles of O_2 by the conversion factor that relates mass and moles, molar mass.

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

Subtract the mass of O_2 needed from the mass available to calculate excess O_2.

$$50.0\text{ g } O_2 \text{ available} - 32.3\text{ g } O_2 \text{ needed} = 17.7\text{ g } O_2 \text{ in excess}$$

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 oxygen that was consumed (32.3 g) and the given mass of phosphorus (25.0 g) is 57.3 g, the calculated mass of the product phosphorus decoxide.

PRACTICE PROBLEMS

20. The reaction between solid sodium and iron(III) oxide is one in a series of reactions that inflates an automobile airbag.

$$6Na(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s)$$

If 100.0 g Na and 100.0 g Fe_2O_3 are used in this reaction, determine

a. the limiting reactant.

b. the reactant in excess.

c. the mass of solid iron produced.

d. the mass of excess reactant that remains after the reaction is complete.

21. Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. Write the balanced chemical equation for the reaction. If a plant has 88.0 g carbon dioxide and 64.0 g water available for photosynthesis, determine

a. the limiting reactant.

b. the excess reactant and the mass in excess.

c. the mass of glucose produced.

For more practice determining the limiting reactant in a chemical reaction, go to **Supplemental Practice Problems** in Appendix A.

Why use an excess of a reactant? Why are reactions usually not carried out using amounts of reactants in the exact mole ratios given in the balanced equation? Some reactions do not continue until all the reactants are used up. Instead, they appear to 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.

In **Figure 12-7,** you can see an example of how controlling the amount of a reactant can increase efficiency. Your school laboratory may have the kind of Bunsen burner shown in the figure. If so, you probably know that this type of burner has a control that can vary the amount of air (oxygen) that mixes with the gas. How efficiently the burner operates depends upon the ratio of oxygen to methane gas in the fuel mixture. When the amount of air is limited, the resulting flame is yellow because of glowing bits of unburned fuel, which deposit on glassware as soot (carbon). 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.

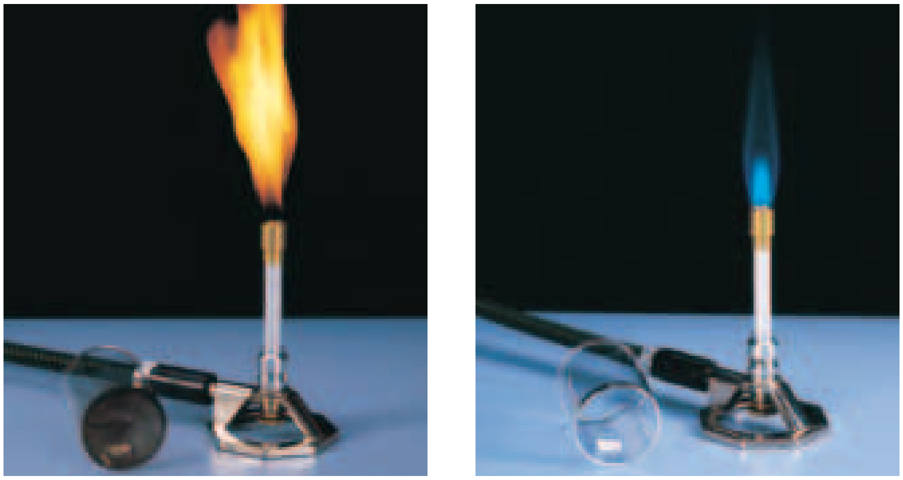


Figure 12-7

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.

Section 12.3 Assessment

22. What is meant by the limiting reactant? Why is it necessary to identify the limiting reactant when you want to know how much product will form in a chemical reaction?

23. Describe how the mass of the product can be calculated when one reactant is in excess.

24. Are limiting reactants present in all reactions? Explain.

25. **Thinking Critically** For the following reactions, identify the limiting reactant and the excess reactant. Give reasons for your choices.

a. wood burning in a campfire

b. sulfur in the air reacting with silver flatware to produce tarnish, or silver sulfide

c. baking powder in cake batter decomposing to produce carbon dioxide, which makes the cake rise

26. **Analyze and Conclude** The equation representing the production of tetraphosphorus trisulfide (P_4S_3), a substance used in some match heads, is

$$8P_4 + 3S_8 \longrightarrow 8P_4S_3$$

Determine if each of the following statements is correct. If the statement is incorrect, rewrite it to make it correct.

a. To produce 4 mol P_4S_3, 4 mol P_4 must react with 1.5 mol S_8.

b. When 4 mol P_4 reacts with 4 mol S_8, sulfur is the limiting reactant.

c. When 6 mol P_4 and 6 mol S_8 react, 1320 g P_4S_3 is produced.

Section 12.4

Percent Yield

Objectives

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

Vocabulary

theoretical yield
actual yield
percent yield

Suppose you were determined to improve your jump shot and took time each afternoon to practice. One afternoon, you succeeded in getting the ball through the hoop 49 times out of a total of 75 tries. Theoretically, you could have been successful 75 times, but in actuality that usually doesn't happen. (That's why you practice.) But how successful were you? You could calculate your efficiency as a percent by dividing the number of successful tries by the total number of tries and multiplying by 100.

$$\frac{49 \text{ actual successes}}{75 \text{ theoretical successes}} \times 100 = 65\% \text{ successful shots}$$

Sixty-five percent successful jump shots means that you could expect to get the ball into the basket 65 times if you made 100 attempts.

Similar calculations are made to determine the success of chemical reactions because most reactions never succeed in producing the predicted amount of product. Although your work with stoichiometric problems so far may have led you to think that chemical reactions proceed according to the balanced equation without any difficulties and always produce the calculated amount of product, this is not the case! Not every reaction goes cleanly or completely. Many reactions stop before all of the reactants are used up, so the actual amount of product is less than expected. Liquid reactants or products may adhere to the surfaces of containers or evaporate, and solid product is always left behind on filter paper or lost in the purification process. In some instances, products other than the intended ones may be formed by competing reactions, thus reducing the yield of the desired product.

Math Handbook

Review the meaning and calculation of percents in the **Math Handbook** on page 907 in this text.

How much product?

In many of the calculations you have been practicing, you have been asked to calculate the amount of product that can be produced from a given amount of reactant. The answer you obtained is called 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 actually produced when the chemical reaction is carried out in an experiment.

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. Just as you calculated your percent of successful jump shots, a chemist can calculate what percent of the amount of product that could theoretically be produced was actually produced. **Percent yield** of product is the ratio of the actual yield to the theoretical yield expressed as a percent.

$$\text{Percent yield} = \frac{\text{actual yield (from an experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100$$

The **problem-solving LAB** on page 372 will help you understand the importance of percent yield in chemical reactions and the kind of factors that may determine the size of the percent yield.

EXAMPLE PROBLEM 12-6

Calculating Percent Yield

When potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g silver nitrate ($AgNO_3$), solid silver chromate (Ag_2CrO_4) is formed.

a. Determine the theoretical yield of the silver chromate precipitate.

b. If 0.455 g of silver chromate is obtained, calculate the percent yield.

1. Analyze the Problem

You are given the mass of the reactant $AgNO_3$ and the actual yield of the product Ag_2CrO_4. You need to write the balanced chemical equation and calculate the theoretical yield by making these conversions: grams of silver nitrate to moles of silver nitrate, moles of silver nitrate to moles of silver chromate, moles of silver chromate to grams of silver chromate. The percent yield can be calculated from the actual yield of product and the calculated 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

Write the balanced chemical equation and indicate the known and unknown quantities.

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

Convert grams of $AgNO_3$ to moles of $AgNO_3$ using the inverse of molar mass.

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

Use the appropriate mole ratio to convert mol $AgNO_3$ to mol Ag_2CrO_4.

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

Calculate the mass of Ag_2CrO_4 (the theoretical yield) by multiplying mol Ag_2CrO_4 by the molar mass.

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

Divide the actual yield by the theoretical yield and multiply by 100.

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

3. Evaluate the Answer

All quantities have three significant figures 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 mol $AgNO_3$ to mol Ag_2CrO_4 in the equation is 2:1. Therefore, 0.500 g $AgNO_3$ 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.

Like silver chromate, all chromium compounds are colored. Although silver chromate is insoluble in water, potassium chromate (K_2CrO_4, yellow) and potassium dichromate ($K_2Cr_2O_7$, orange) are soluble.

For more practice calculating percent yield, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

27. Aluminum hydroxide is often present in antacids to neutralize stomach acid (HCl). If 14.0 g aluminum hydroxide is present in an antacid tablet, determine the theoretical yield of aluminum chloride produced when the tablet reacts with stomach acid. If the actual yield of aluminum chloride from this tablet is 22.0 g, what is the percent yield?

$Al(OH)_3(s) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$

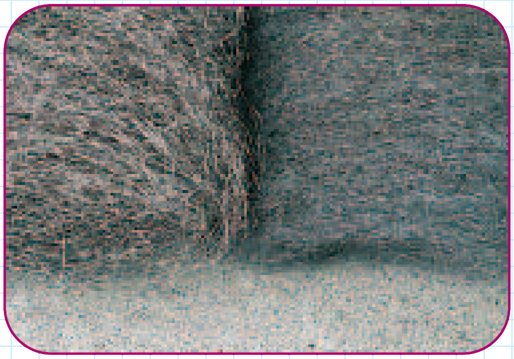
28. When copper wire is placed into a silver nitrate solution, silver crystals and copper(II) nitrate solution form. Write the balanced chemical equation for the reaction. If a 20.0-g sample of copper is used, determine the theoretical yield of silver. If 60.0 g silver is actually recovered from the reaction, determine the percent yield of the reaction.

29. Zinc reacts with iodine in a synthesis reaction. Write the balanced chemical equation for the reaction. Determine the theoretical yield if a 125.0-g sample of zinc was used. Determine the percent yield if 515.6 g product is recovered.

problem-solving LAB

How does the surface area of a solid reactant affect percent yield?

Designing an Experiment The cost of every manufactured item you buy is based largely on the cost of producing the item. Manufacturers compete to reduce costs and increase profits. This means increasing the percent yield of the manufacturing process by producing the most product for the amount of reactant used. If you were going to produce iron oxide (Fe_2O_3) from steel wool, how could you design an experiment to determine what gauge (diameter) steel wool will produce the highest yield? Write the equation for the reaction upon which you will base your experiment.



Analysis

In the first photo, different gauges of steel wool are shown. The second photo shows the combustion of a sample of steel wool using a Bunsen burner.

Thinking Critically

1. What quantities must be used to calculate the percent yield of Fe_2O_3 when iron is burned? How will these quantities be measured and how many measurements should be made?
2. What quantities should be kept constant in the experiment?
3. How will the resulting data be analyzed?
4. Are there any obvious errors in the design that could significantly affect the results? If so, how could they be avoided?

Percent yield in the marketplace You learned in the **problem-solving LAB** that in order to compete, manufacturers must reduce the cost of making their products to the lowest level possible. Percent yield is important in the calculation of overall cost effectiveness in industrial processes. For example, sulfuric acid (H_2SO_4) is made using mined sulfur, **Figure 12-8.** Sulfuric acid is an important chemical because it is a raw material for 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.

A two-step process called the contact process is often used for the manufacture of sulfuric acid. Over time, the process has been improved by chemical engineers to produce the maximum yield of product and, at the same time, comply with environmental standards for clean air. The two steps in the contact process are

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

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

A final step, the combination of SO_3 with water, produces the product, H_2SO_4.

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

The first step, the combustion of sulfur, produces 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 speeds up the reaction but the yield decreases.

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%. This four-pass procedure maximizes the yield at temperatures near 400°C, and uses the modest increase in temperature to increase the rate and minimize the time.

Figure 12-8

The large tonnage of sulfur needed to make sulfuric acid and other products is often obtained by forcing hot water into underground deposits to melt the sulfur and then pumping the liquid sulfur to the surface.

Section 12.4 Assessment

30. Distinguish between the theoretical yield and the actual yield of a chemical reaction.

31. Give several reasons why the actual yield is not usually equal to the theoretical yield.

32. Explain how percent yield is calculated.

33. **Thinking Critically** In an experiment, you are to combine iron with an excess of sulfur and heat the mixture to obtain iron(III) sulfide.

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

What experimental information must you collect in order to calculate the percent yield of this reaction?

34. **Interpreting Data** Use the data to determine the percent yield of the following reaction.
$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
Oxygen is in excess.

Reaction Data

Mass of crucible	35.67 g
Mass of crucible + Mg	38.06 g
Mass of Mg	
Mass of crucible + MgO	39.15 g
Mass of MgO	

CHEMLAB 12

A Mole Ratio

Iron reacts with copper(II) sulfate in a single replacement reaction. By measuring the mass of iron that reacts and the mass of copper metal produced, you can calculate the ratio of moles of reactant to moles of product. This mole ratio can be compared to the ratio found in the balanced chemical equation.

Problem

Which reactant is the limiting reactant? How does the experimental mole ratio of Fe to Cu compare with the mole ratio in the balanced chemical equation? What is the percent yield?

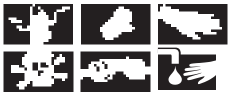
Objectives

- **Observe** a single replacement reaction.
- **Measure** the masses of iron and copper.
- **Calculate** the moles of each metal and the mole ratio.

Materials

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

Safety Precautions



- **Always wear safety glasses and a lab apron.**
- **Hot objects will not appear to be hot.**
- **Do not heat broken, chipped, or cracked glassware.**
- **Turn off the hot plate when not in use.**

Pre-Lab

1. Read the entire **CHEMLAB**.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Is it important that you know you are using the hydrated form of copper(II) sulfate? Would it be possible to use the anhydrous form? Why or why not?

Data for the Reaction of Copper(II) Sulfate and Iron	
Mass of empty 150-mL beaker	
Mass of 150-mL beaker + $CuSO_4 \cdot 5H_2O$	
Mass of $CuSO_4 \cdot 5H_2O$	
Mass of iron filings	
Mass of 150-mL beaker and dried copper	
Mass of dried copper	
Observations	

Procedure

1. Measure and record the mass of a clean, dry 150-mL beaker.
2. Place approximately 12 g of copper(II) sulfate pentahydrate into the 150-mL beaker and measure and record the combined mass.
3. Add 50 mL of distilled water to the copper(II) sulfate pentahydrate and heat the mixture on the hot plate at a medium setting. Stir until all of the solid is dissolved, but do not boil. Using tongs, remove the beaker from the hot plate.
4. Measure approximately 2 g of iron metal filings onto a piece of weighing paper. Measure and record the exact mass of the filings.
5. While stirring, slowly add the iron filings to the hot copper(II) sulfate solution.
6. Allow the reaction mixture to stand, without stirring, for five minutes to ensure complete reaction. The solid copper metal will settle to the bottom of the beaker.
7. Use the stirring rod to decant (pour off) the liquid into a 400-mL beaker. Be careful to decant only the liquid.

8. 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.
9. Repeat step 8 two more times.
10. Place the 150-mL beaker containing the wet copper on the hot plate. Use low heat to dry the copper.
11. Remove the beaker from the hot plate and allow it to cool.
12. Measure and record the mass of the cooled 150-mL beaker and the copper.

Cleanup and Disposal

1. Make sure the hot plate is off.
2. The dry copper can be placed in a waste container. Wet 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 in the fume hood.
3. Return all lab equipment to its proper place.
4. Wash your hands thoroughly after all lab work and cleanup is complete.

Analyze and Conclude

1. **Observing and Inferring** What evidence did you observe that confirms that a chemical reaction occurred?
2. **Applying Concepts** Write a balanced chemical equation for the single-replacement reaction that occurred.
3. **Interpreting Data** From your data, determine the mass of copper produced.
4. **Using Numbers** Use the mass of copper to calculate the moles of copper produced.
5. **Using Numbers** Calculate the moles of iron used in the reaction.
6. **Using Numbers** Determine the whole number ratio of moles of iron to moles of copper.
7. **Comparing and Contrasting** Compare the ratio of moles of iron to moles of copper from the balanced chemical equation to the mole ratio calculated using your data.
8. **Evaluating Results** Use the balanced chemical equation to calculate the mass of copper that should have been produced from the sample of iron you used. Use this number and the mass of copper you actually obtained to calculate the percent yield.
9. **Error Analysis** What was the source of any deviation from the mole ratio calculated from the chemical equation? How could you improve your results?
10. **Drawing a Conclusion** Which reactant is the limiting reactant? Explain.

Real-World Chemistry

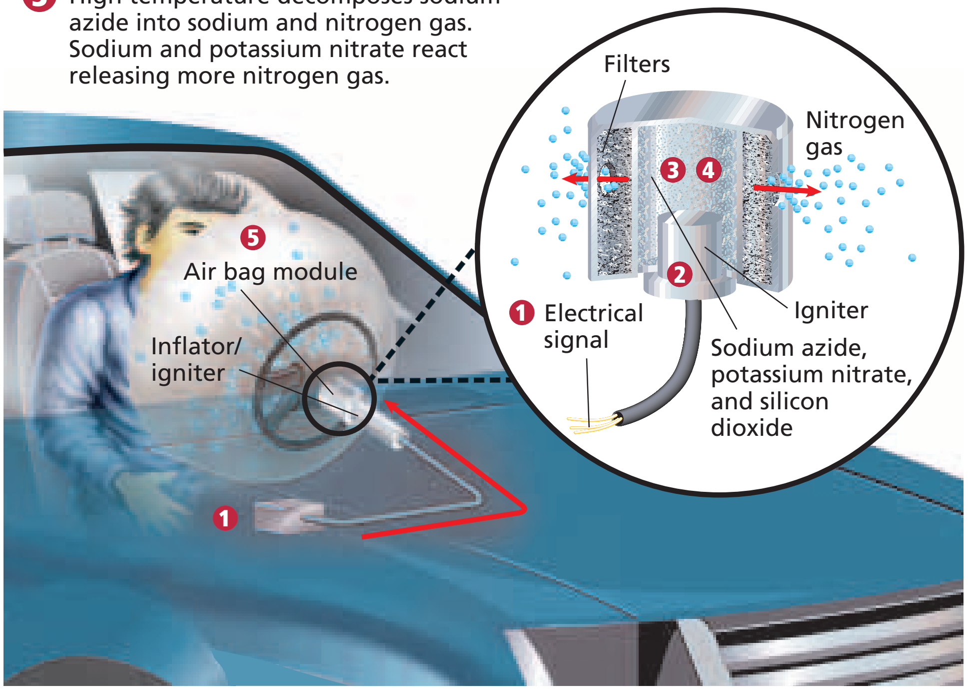
1. A furnace that provides heat by burning methane gas (CH_4) must have the correct mixture of air and fuel to operate efficiently. What is the mole ratio of air to methane gas in the combustion of methane? Hint: Air is 20% oxygen.
2. Automobile air bags inflate on impact because a series of gas-producing chemical reactions are triggered. To be effective in saving lives, the bags must not overinflate or underinflate. What factors must automotive engineers take into account in the design of air bags?

How It Works

Air Bags

Air bags fill with nitrogen gas as they deploy in automobile crashes. The source of the nitrogen gas is the chemical compound sodium azide (NaN_3). Hazardous sodium metal is produced along with the nitrogen, so potassium nitrate is added to convert the sodium into less hazardous sodium oxide (Na_2O). Stoichiometric calculations are needed to determine the precise quantity of sodium azide that will produce the volume of nitrogen gas required to inflate the air bag. If too much gas is produced, the air bag may be so rigid that hitting it would be the same as hitting a solid wall.

1. Crash sensor detects rapid deceleration and sends signal to air bag module.
2. Igniter explodes, heating sodium azide in the inflator.
3. High temperature decomposes sodium azide into sodium and nitrogen gas. Sodium and potassium nitrate react releasing more nitrogen gas.
4. Silicon dioxide (sand), sodium oxide, and potassium oxide formed in the inflator fuse into glass.
5. Nitrogen gas inflates air bag.



Thinking Critically

1. **Predicting** Which starting material used in the air bag inflator is the least important for the proper inflation of the air bag? Would it be necessary to have it present in a precise stoichiometric ratio? Why or why not?

2. **Analyze and Conclude** What is the correct stoichiometric ratio between NaN_3 and KNO_3 to ensure no sodium is unreacted? What would be the consequences of an excess of KNO_3 to the operation of an air bag?

Summary

12.1 What is stoichiometry?

- Balanced chemical equations can be interpreted in terms of representative particles (atoms, molecules, formula units), moles, and mass.
- The law of conservation of mass, as applied to chemical reactions, means that the total mass of the reactants is equal to the total mass of the products.
- Mole ratios are central to stoichiometric calculations. They are derived from the coefficients in a balanced chemical equation. To write mole ratios, the number of moles of each reactant and product is placed, in turn, in the numerator of the ratio with the moles of each other reactant and product placed in the denominator.

12.2 Stoichiometric Calculations

- Stoichiometric calculations allow a chemist to predict the amount of product that can be obtained from a given amount of reactant or to determine how much of two or more reactants must be used to produce a specified amount of product.
- The four steps in stoichiometric calculations begin with the balanced chemical equation.
- Mole ratios used in the calculations are determined from the balanced chemical equation.
- The mass of the given substance is converted to moles of the given substance. Then, moles of the given substance are converted by means of a mole ratio to moles of the unknown substance. Finally, moles of the unknown substance are converted to the mass of the unknown substance.

12.3 Limiting Reactants

- 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 given amount of the limiting reactant.

12.4 Percent Yield

- 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 actually 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.

Key Equations and Relationships

- $\text{moles of known} \times \dfrac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$
  (p. 358)

- $\dfrac{\text{actual yield (from experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100 = \text{percent yield}$
  (p. 370)

Vocabulary

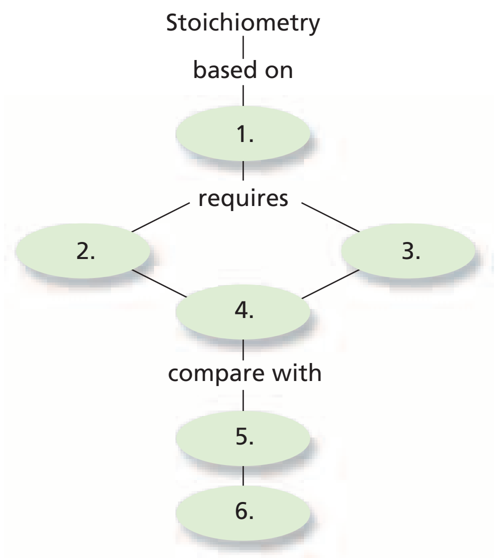
- actual yield (p. 370)
- excess reactant (p. 364)
- limiting reactant (p. 364)
- mole ratio (p. 356)
- percent yield (p. 370)
- stoichiometry (p. 354)
- theoretical yield (p. 370)

chemistrymc.com/vocabulary_puzzlemaker

CHAPTER 12 ASSESSMENT

Go to the Chemistry Web site at **chemistrymc.com** *for additional Chapter 12 Assessment.*

Concept Mapping



35. Fill in the ovals with the following terms to create a concept map: actual yield, balanced chemical equation, molar mass, mole ratio, percent yield, and theoretical yield.

Mastering Concepts

36. What relationships can be determined from a balanced chemical equation? (12.1)

37. Explain how the law of conservation of mass allows you to interpret a balanced chemical equation in terms of mass. (12.1)

38. Explain why mole ratios are central to stoichiometric calculations. (12.1)

39. What is the mole ratio that can convert from moles of A to moles of B? (12.1)

40. What is the first step in all stoichiometric calculations? (12.2)

41. How is molar mass used in some stoichiometric calculations? (12.2)

42. What information must you have in order to calculate the mass of product formed in a chemical reaction? (12.2)

43. What is meant by limiting reactant? Excess reactant? (12.3)

44. How are mole ratios used in finding the limiting reactant in a reaction? (12.3)

45. What is the difference between actual yield and theoretical yield? (12.4)

46. How are actual yield and theoretical yield determined? (12.4)

47. Can the percent yield of a chemical reaction be more than 100%? Explain your answer. (12.4)

48. What relationship is used to determine the percent yield of a chemical reaction? (12.4)

49. What experimental information do you need in order to calculate both the theoretical and percent yield of any chemical reaction? (12.4)

50. 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? (12.4)

Mastering Problems

Interpreting Equations (12.1)

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

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

52. 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 equation in terms of particles, moles, and mass.

53. 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.

Mole Ratios (12.1)

54. 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.

55. 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 mol Fe if mol Fe_2O_3 is known?

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

56. 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.

57. Determine the mole ratio necessary to convert moles of aluminum to moles of aluminum chloride when aluminum reacts with chlorine.

58. Chromite ($FeCr_2O_4$) is the most important commercial ore of chromium. 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?

59. The air pollutant SO_2 is removed from the air by means of a reaction among sulfur dioxide, calcium carbonate, and oxygen. The products of this reaction are calcium sulfate and carbon dioxide. Determine the mole ratio you would use to convert mol SO_2 to mol $CaSO_4$.

Reaction Data

Moles of reactants		Moles of products	
W	X	Y	Z
0.90	0.30	0.60	1.20

60. Two substances, W and X, react to form the products Y and Z. The table shows the numbers of moles of the reactants and products involved when the reaction was carried out in one experiment. Use the data to determine the coefficients that will balance the equation $W + X \longrightarrow Y + Z$.

Stoichiometric Mole-to-Mole Conversions (12.2)

61. If 5.50 mol calcium carbide (CaC_2) reacts with an excess of water, how many moles of acetylene (C_2H_2) will be produced?

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

62. When an antacid tablet dissolves in water, the fizz is due to a reaction between sodium hydrogen carbonate (sodium bicarbonate, $NaHCO_3$) 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 carbon dioxide can be produced if one tablet containing 0.0119 mol $NaHCO_3$ is dissolved?

63. One of the main components of pearls is calcium carbonate. If pearls are put in acidic solution, they dissolve.

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

How many mol $CaCO_3$ can be dissolved in 0.0250 mol HCl?

Stoichiometric Mole-to-Mass Conversions (12.2)

64. Citric acid ($H_3C_6H_5O_7$) is a product of the fermentation of sucrose ($C_{12}H_{22}O_{11}$) in air.

$C_{12}H_{22}O_{11}(aq) + 3O_2(g) \longrightarrow 2H_3C_6H_5O_7(aq) + 3H_2O(l)$

Determine the mass of citric acid produced when 2.50 mol $C_{12}H_{22}O_{11}$ is used.

65. Esterification is a reaction between an organic acid and an alcohol that forms as ester and water. The ester ethyl butanoate ($C_3H_7COOC_2H_5$), which is responsible for the fragrance of pineapples, 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 ethanol is used.

66. Carbon dioxide 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 CO_2.

67. 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 potassium chromate, determine the mass of lead chromate that can be obtained.

68. 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 many grams of hydrazine are needed to produce 10.0 mol nitrogen gas?

Stoichiometric Mass-to-Mass Conversions (12.2)

69. 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 many g CH_4 is needed to produce 50.0 g $CHCl_3$?

70. 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 ethanol.

____$C_2H_5OH(l)$ + ____$O_2(g) \rightarrow$____$CO_2(g)$ + ____$H_2O(g)$

71. When surface water dissolves carbon dioxide, carbonic acid (H_2CO_3) is formed. When the water moves underground through limestone formations, the limestone dissolves and caves are sometimes produced.

$CaCO_3(s) + H_2CO_3(aq) \longrightarrow Ca(HCO_3)_2(aq)$

What mass of limestone must have dissolved if 3.05×10^{10} kg of calcium hydrogen carbonate was produced?

72. Car batteries use solid lead and lead(IV) oxide with sulfuric acid solution to produce an electric current. The products of this reaction are lead(II) sulfate in solution and water.

a. Write the balanced chemical equation for this reaction.
b. Determine the mass of lead(II) sulfate produced when 25.0 g lead reacts with an excess of lead(IV) oxide and sulfuric acid.

73. The fuel methanol (CH_3OH) is made by the reaction of carbon monoxide and hydrogen.

a. Write the balanced chemical equation.
b. How many grams of hydrogen are needed to produce 45.0 grams of methanol?

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

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

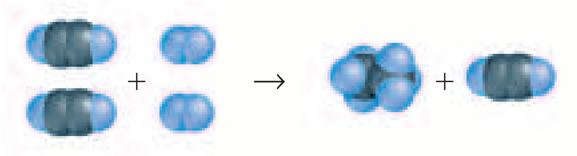
a. Determine the mass of gold that can be extracted if 25.0 g sodium cyanide is used.
b. If the mass of the ore from which the gold was extracted is 150.0 g, what percentage of the ore is gold?

75. 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 AgBr is removed.

Limiting Reactants (12.3)



76. The illustration shows the reaction between ethyne (acetylene, C_2H_2) and hydrogen. The product is ethane (C_2H_6). Which is the limiting reactant? Which is the excess reactant? Explain.

77. This reaction takes place in a nickel-iron battery.

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

Determine the number of moles of iron(II) hydroxide ($Fe(OH)_2$) produced if 5.00 mol Fe and 8.00 mol NiO(OH) react.

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

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

79. Iron is obtained commercially by the reaction of hematite (Fe_2O_3) with carbon monoxide. How many grams of iron are produced if 25.0 moles of hematite react with 30.0 moles of carbon monoxide?

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

80. Under certain conditions of temperature and pressure, hydrogen and nitrogen react to produce ammonia (NH_3). Write the balanced chemical equation and determine the mass of ammonia produced if 3.50 mol H_2 reacts with 5.00 mol N_2.

81. The reaction of chlorine gas with solid phosphorus (P_4) produces solid phosphorus pentachloride. When 16.0 g chlorine reacts with 23.0 g P_4, which reactant limits the amount of phosphorus pentachloride produced? Which reactant is in excess?

82. 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 Zn and 30.0 g MnO_2 are used.
b. Determine the mass of $Zn(OH)_2$ produced.

83. 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 mass in excess.

Percent Yield (12.4)

84. Ethanol (C_2H_5OH) is produced from the fermentation of sucrose 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 and percent yields of ethanol if 684 g sucrose undergoes fermentation and 349 g ethanol is obtained.

85. Lead(II) oxide is obtained by roasting galena, lead(II) sulfide, in air.

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

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

86. Upon heating, calcium carbonate decomposes to produce calcium oxide and carbon dioxide.

a. Determine the theoretical yield of CO_2 if 235.0 g $CaCO_3$ is heated.
b. What is the percent yield of CO_2 if 97.5 g CO_2 is collected?

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

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

If 40.0 g SiO_2 and 40.0 g of HF react

a. determine the limiting reactant.
b. determine the mass of the excess reactant.
c. determine the theoretical yield of H_2SiF_6.
d. determine the percent yield if the actual yield is 45.8 g H_2SiF_6.

88. 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 ZrI_4 is decomposed and 5.00 g pure Zr is obtained.

89. Phosphorus is commercially prepared by heating a mixture of calcium phosphate, sand, and coke 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 $Ca_3(PO_4)_2$ and 400.0 g SiO_2 are heated. If the actual yield of P_4 is 45.0 g, determine the percent yield of P_4.

90. Chlorine can be prepared by the reaction of manganese(IV) oxide with hydrochloric acid.

____ $MnO_2(s)$ + ____ $HCl(aq) \longrightarrow$

____ $MnCl_2(aq)$ + ____ $Cl_2(g)$ + ____ $H_2O(l)$

Balance the equation and determine the theoretical and percent yields of chlorine if 86.0 g MnO_2 and 50.0 g HCl react. The actual yield of chlorine is 20.0 g.

Mixed Review

Sharpen your problem-solving skills by answering the following.

91. 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?

92. One method for producing nitrogen in the laboratory is to react ammonia with copper(II) oxide.

____ $NH_3(g)$ + ____ $CuO(s) \longrightarrow$

____ $Cu(s)$ + ____ $H_2O(l)$ + ____ $N_2(g)$

a. Balance the equation.
b. If 40.0 g NH_3 is reacted with 80.0 g CuO, determine the limiting reactant.
c. Determine the mass of N_2 produced by this reaction.
d. Which reactant is in excess? How much remains after the reaction?

93. The compound calcium cyanamide (CaNCN) can be used as a fertilizer. 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 CaC_2 reacts with 5.00 mol N_2?

94. 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 copper(II) oxide is used?

95. Nitrogen oxide is present in urban pollution but it is immediately converted 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?

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

97. The Swedish chemist Karl Wilhellm was first to produce chlorine in the laboratory.

$2NaCl(s) + 2H_2SO_4(aq) + MnO_2(aq) \rightarrow Na_2SO_4(aq) + MnSO_4(aq) + 2H_2O(l) + Cl_2(g).$

What mole ratio could be used to find the moles of chlorine produced from 4.85 moles of sodium chloride? Determine the moles of chlorine produced. Determine the mass of chlorine produced.

98. The solid booster rockets of the space shuttle contain ammonium perchlorate (NH_4ClO_4) and powdered aluminum as the propellant.

$8Al + 3NH_4ClO_4 \rightarrow 4Al_2O_3 + 3NH_4Cl.$

Determine the percent yield if 6.00×10^5 kg NH_4ClO_4 produces 6.56×10^5 kg aluminum oxide.

Thinking Critically

99. 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?

100. Observing and Inferring Determine whether the following reactions depend upon 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.
c. Propane (C_3H_8) burns in excess oxygen to produce carbon dioxide and water.

101. Designing 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.

102. Formulating Models Copper reacts with chlorine to produce copper(II) chloride. Draw a diagram that represents eight atoms of copper reacting with six molecules of chlorine. Make sure you include the particles before the reaction and after the reaction. Include any excess reactants.

103. Applying Concepts When your campfire begins to die down and smolder, it helps to fan it. Explain in terms of stoichiometry why the fire begins to flare up again.

Writing in Chemistry

104. Research the air pollutants produced by using 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.

105. 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.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

106. 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 theory and why? (Chapter 1)

107. Write the electron configuration for each of the following atoms. (Chapter 5)

a. fluorine
b. aluminum
c. titanium
d. radon

108. Explain why the gaseous nonmetals exist as diatomic molecules, but other gaseous elements exist as single atoms. (Chapter 9)

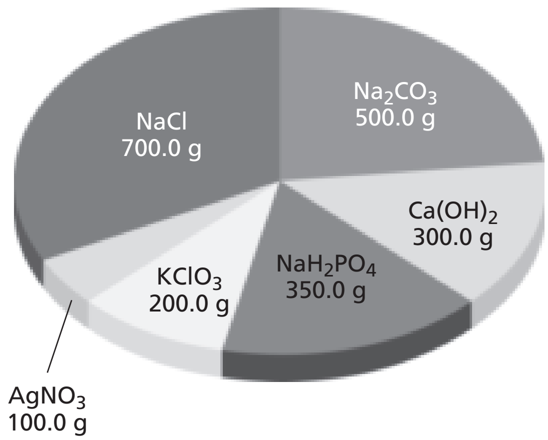
109. Write a balanced equation for the reaction of potassium with oxygen. (Chapter 10)

110. What is the molecular mass of UF_6? What is the molar mass of UF_6? (Chapter 11)

Use these questions and the test-taking tip to prepare for your standardized test.

Interpreting Graphs Use the graph below to answer questions 1–4.

Supply of Various Chemicals in Dr. Raitano's Laboratory



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

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

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 **c.** 74.7 g
b. 37.3 g **d.** 100 g

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

The LeBlanc process, shown above, is the traditional method of manufacturing sodium hydroxide. 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 **c.** 8.097 mol
b. 4.72 mol **d.** 9.43 mol

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

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

If half of the $KClO_3$ in the lab is used and 12.8 g of oxygen gas are produced, what is the percent yield of this reaction?

a. 12.8% **c.** 65.6%
b. 32.7% **d.** 98.0%

4. Sodium dihydrogen pyrophosphate ($Na_2H_2P_2O_7$), more commonly known as baking powder, is manufactured by heating NaH_2PO_4 at high temperatures:

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

If 444.0 g of $Na_2H_2P_2O_7$ are needed, how much more NaH_2PO_4 will Dr. Raitano have to buy to make enough $Na_2H_2P_2O_7$?

a. 94.0 g
b. 130.0 g
c. 480 g
d. none—the lab already has enough

5. Stoichiometry is based on the law of

a. constant mole ratios.
b. Avogadro's constant.
c. conservation of energy.
d. conservation of mass.

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

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

If 3.55 moles of HgO decompose to form 1.54 moles of O_2 and 618 g of Hg, what is the percent yield of this reaction?

a. 13.2% **c.** 56.6%
b. 42.5% **d.** 86.8%

7. Dimethyl hydrazine $(CH_3)_2N_2H_2$ ignites spontaneously upon 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 2.0 moles of dimethyl hydrazine are mixed with 4.0 moles of dinitrogen tetroxide, and the reaction achieves an 85% yield, how many moles of N_2, H_2O, and CO_2 will be formed?

a. 0.57 mol N_2, 0.43 mol H_2O, 0.85 mol CO_2
b. 2.6 mol N_2, 3.4 mol H_2O, 1.7 mol CO_2
c. 5.1 mol N_2, 6.8 mol H_2O, 3.4 mol CO_2
d. 6.0 mol N_2, 8.0 mol H_2O, 4.0 mol CO_2

TEST-TAKING TIP

Calculators Are Only Machines If your test allows you to use a calculator, use it wisely. The calculator can't figure out what the question is asking. That's still *your* job. Figure out which numbers are relevant, and determine the best way to solve the problem before you start punching keys.