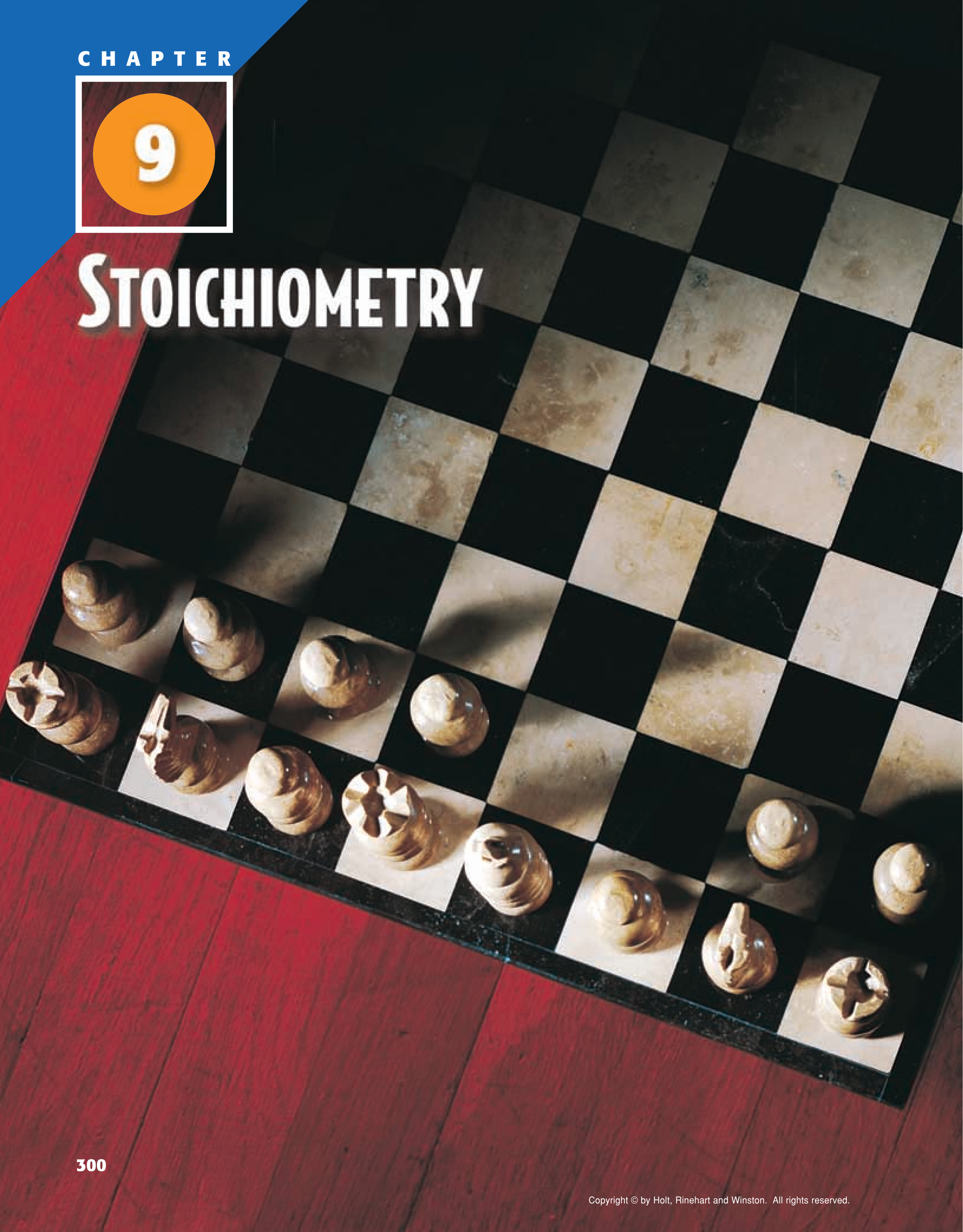

CHAPTER 9

STOICHIOMETRY

To play a standard game of chess, each side needs the proper number of pieces and pawns. Unless you find all of them—a king, a queen, two bishops, two knights, two rooks, and eight pawns—you cannot start the game. In chemical reactions, if you do not have every reactant, you will not be able to start the reaction. In this chapter you will look at amounts of reactants present and calculate the amounts of other reactants or products that are involved in the reaction.

START-UP ACTIVITY

All Used Up

PROCEDURE

1. Use a **balance** to find the mass of **8 nuts** and the mass of **5 bolts**.
2. Attach 1 nut (N) to 1 bolt (B) to assemble a nut-bolt (NB) model. Make as many NB models as you can. Record the number of models formed, and record which material was used up. Take the models apart.
3. Attach 2 nuts to 1 bolt to assemble a nut-nut-bolt (N_2B) model. Make as many N_2B models as you can. Record the number of models formed, and record which material was used up. Take the models apart.

ANALYSIS

1. Using the masses of the starting materials (the nuts and the bolts), could you predict which material would be used up first? Explain.
2. Write a balanced equation for the "reaction" that forms NB. How can this equation help you predict which component runs out?
3. Write a balanced equation for the "reaction" that forms N_2B. How can this equation help you predict which component runs out?
4. If you have 18 bolts and 26 nuts, how many models of NB could you make? of N_2B?

Pre-Reading Questions

1. **A recipe calls for one cup of milk and three eggs per serving. You quadruple the recipe because you're expecting guests. How much milk and eggs do you need?**
2. **A bicycle mechanic has 10 frames and 16 wheels in the shop. How many complete bicycles can he assemble using these parts?**
3. **List at least two conversion factors that relate to the mole.**

CONTENTS 9

SECTION

1

Calculating Quantities in Reactions

KEY TERMS

- **stoichiometry**

OBJECTIVES

1. **Use** proportional reasoning to determine mole ratios from a balanced chemical equation.
2. **Explain** why mole ratios are central to solving stoichiometry problems.
3. **Solve** stoichiometry problems involving mass by using molar mass.
4. **Solve** stoichiometry problems involving the volume of a substance by using density.
5. **Solve** stoichiometry problems involving the number of particles of a substance by using Avogadro's number.

Balanced Equations Show Proportions

If you wanted homemade muffins, like the ones in **Figure 1,** you could make them yourself—if you had the right things. A recipe for muffins shows how much of each ingredient you need to make 12 muffins. It also shows the proportions of those ingredients. If you had just a little flour on hand, you could determine how much of the other things you should use to make great muffins. The proportions also let you adjust the amounts to make enough muffins for all your classmates.

A balanced chemical equation is very similar to a recipe in that the coefficients in the balanced equation show the proportions of the reactants and products involved in the reaction. For example, consider the reaction for the synthesis of water.

$$2H_2 + O_2 \rightarrow 2H_2O$$

On a very small scale, the coefficients in a balanced equation represent the numbers of particles for each substance in the reaction. For the equation above, the coefficients show that two molecules of hydrogen react with one molecule of oxygen and form two molecules of water.

Calculations that involve chemical reactions use the proportions from balanced chemical equations to find the quantity of each reactant and product involved. As you learn how to do these calculations in this section, you will assume that each reaction goes to completion. In other words, all of the given reactant changes into product. For each problem in this section, assume that there is more than enough of all other reactants to completely react with the reactant given. Also assume that every reaction happens perfectly, so that no product is lost during collection. As you will learn in the next section, this usually is not the case.

Figure 1
In using a recipe to make muffins, you are using proportions to determine how much of each ingredient is needed.

Relative Amounts in Equations Can Be Expressed in Moles

Just as you can interpret equations in terms of particles, you can interpret them in terms of moles. The coefficients in a balanced equation also represent the moles of each substance. For example, the equation for the synthesis of water shows that 2 mol H_2 react with 1 mol O_2 to form 2 mol H_2O. Look at the equation below.

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

This equation shows that 2 molecules C_8H_{18} react with 25 molecules O_2 to form 16 molecules CO_2 and 18 molecules H_2O. And because Avogadro's number links molecules to moles, the equation also shows that 2 mol C_8H_{18} react with 25 mol O_2 to form 16 mol CO_2 and 18 mol H_2O.

In this chapter you will learn to determine how much of a reactant is needed to produce a given quantity of product, or how much of a product is formed from a given quantity of reactant. The branch of chemistry that deals with quantities of substances in chemical reactions is known as **stoichiometry.**

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stoichiometry

the proportional relationship between two or more substances during a chemical reaction

The Mole Ratio Is the Key

If you normally buy a lunch at school each day, how many times would you need to "brown bag" it if you wanted to save enough money to buy a CD player? To determine the answer, you would use the units of dollars to bridge the gap between a CD player and school lunches. In stoichiometry problems involving equations, the unit that bridges the gap between one substance and another is the mole.

The coefficients in a balanced chemical equation show the relative numbers of moles of the substances in the reaction. As a result, you can use the coefficients in conversion factors called *mole ratios*. Mole ratios bridge the gap and can convert from moles of one substance to moles of another, as shown in **Skills Toolkit 1.**

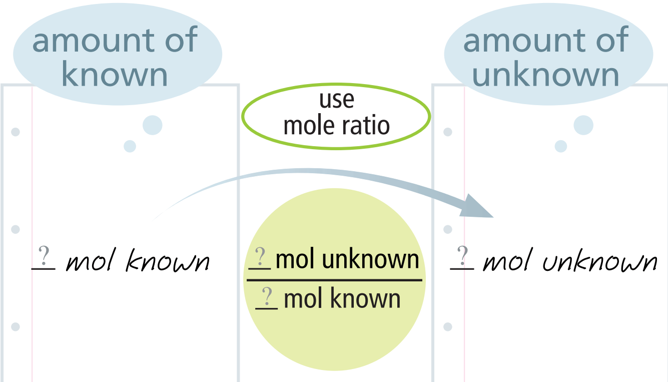
SKILLS Toolkit 1

Converting Between Amounts in Moles

1. Identify the amount in moles that you know from the problem.

2. Using coefficients from the balanced equation, set up the mole ratio with the known substance on bottom and the unknown substance on top.

3. Multiply the original amount by the mole ratio.



SAMPLE PROBLEM A

Using Mole Ratios

Consider the reaction for the commercial preparation of ammonia.

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

How many moles of hydrogen are needed to prepare 312 moles of ammonia?

1 Gather information.

- amount of NH_3 = 312 mol
- amount of H_2 = ? mol
- From the equation: 3 mol H_2 = 2 mol NH_3.

2 Plan your work.

The mole ratio must cancel out the units of mol NH_3 given in the problem and leave the units of mol H_2. Therefore, the mole ratio is

$$\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3}$$

3 Calculate.

$$? \text{ mol } H_2 = 312 \cancel{\text{ mol } NH_3} \times \frac{3 \text{ mol } H_2}{2 \cancel{\text{ mol } NH_3}} =$$

$$468 \text{ mol } H_2$$

4 Verify your result.

- The answer is larger than the initial number of moles of ammonia. This is expected, because the conversion factor is greater than one.
- The number of significant figures is correct because the coefficients 3 and 2 are considered to be exact numbers.

PRACTICE

PROBLEM SOLVING SKILL

1 Calculate the amounts requested if 1.34 mol H_2O_2 completely react according to the following equation.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

a. moles of oxygen formed
b. moles of water formed

2 Calculate the amounts requested if 3.30 mol Fe_2O_3 completely react according to the following equation.

$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

a. moles of aluminum needed
b. moles of iron formed
c. moles of aluminum oxide formed

PRACTICE HINT

The mole ratio must always have the unknown substance on top and the substance given in the problem on bottom for units to cancel correctly.

Getting into Moles and Getting out of Moles

Substances are usually measured by mass or volume. As a result, before using the mole ratio you will often need to convert between the units for mass and volume and the unit *mol*. Yet each stoichiometry problem has the step in which moles of one substance are converted into moles of a second substance using the mole ratio from the balanced chemical equation. Follow the steps in **Skills Toolkit 2** to understand the process of solving stoichiometry problems.

The thought process in solving stoichiometry problems can be broken down into three basic steps. First, change the units you are given into moles. Second, use the mole ratio to determine moles of the desired substance. Third, change out of moles to whatever unit you need for your final answer. And if you are given moles in the problem or need moles as an answer, just skip the first step or the last step! As you continue reading, you will be reminded of the conversion factors that involve moles.

Solving Stoichiometry Problems

You can solve all types of stoichiometry problems by following the steps outlined below.

1. Gather information.

- If an equation is given, make sure the equation is balanced. If no equation is given, write a balanced equation for the reaction described.
- Write the information provided for the given substance. If you are not given an amount in moles, determine the information you need to change the given units into moles and write it down.
- Write the units you are asked to find for the unknown substance. If you are not asked to find an amount in moles, determine the information you need to change moles into the desired units, and write it down.
- Write an equality using substances and their coefficients that shows the relative amounts of the substances from the balanced equation.

2. Plan your work.

- Think through the three basic steps used to solve stoichiometry problems: change to moles, use the mole ratio, and change out of moles. Know which conversion factors you will use in each step.
- Write the mole ratio you will use in the form:

$$\frac{\text{moles of unknown substance}}{\text{moles of given substance}}$$

3. Calculate.

- Write a question mark with the units of the answer followed by an equals sign and the quantity of the given substance.
- Write the conversion factors—including the mole ratio—in order so that you change the units of the given substance to the units needed for the answer.
- Cancel units and check that the remaining units are the required units of the unknown substance.
- When you have finished your calculations, round off the answer to the correct number of significant figures. In the examples in this book, only the final answer is rounded off.
- Report your answer with correct units and with the name or formula of the substance.

4. Verify your result.

- Verify your answer by estimating. You could round off the numbers in the setup in step 3 and make a quick calculation. Or you could compare conversion factors in the setup and decide whether the answer should be bigger or smaller than the initial value.
- Make sure your answer is reasonable. For example, imagine that you calculate that 725 g of a reactant is needed to form 5.3 mg (0.0053 g) of a product. The large difference in these quantities should alert you that there may be an error and that you should double-check your work.

Figure 2
These tanks store ammonia for use as fertilizer. Stoichiometry is used to determine the amount of ammonia that can be made from given amounts of H_2 and N_2.

Problems Involving Mass, Volume, or Particles

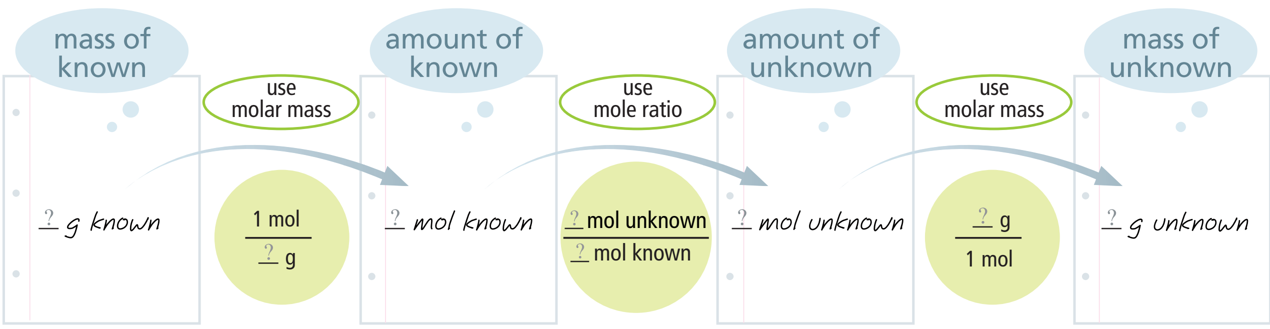
Figure 2 shows a few of the tanks used to store the millons of metric tons of ammonia made each year in the United States. Stoichiometric calculations are used to determine how much of the reactants are needed and how much product is expected. However, the calculations do not start and end with moles. Instead, other units, such as liters or grams, are used. Mass, volume, or number of particles can all be used as the starting and ending quantities of stoichiometry problems. Of course, the key to each of these problems is the mole ratio.

For Mass Calculations, Use Molar Mass

The conversion factor for converting between mass and amount in moles is the molar mass of the substance. The molar mass is the sum of atomic masses from the periodic table for the atoms in a substance. **Skills Toolkit 3** shows how to use the molar mass of each substance involved in a stoichiometry problem. Notice that the problem is a three-step process. The mass in grams of the given substance is converted into moles. Next, the mole ratio is used to convert into moles of the desired substance. Finally, this amount in moles is converted into grams.

3 SKILLS Toolkit

Solving Mass-Mass Problems



SAMPLE PROBLEM B

Problems Involving Mass

What mass of NH_3 can be made from 1221 g H_2 and excess N_2?

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

1 Gather information.

- mass of H_2 = 1221 g H_2
- molar mass of H_2 = 2.02 g/mol
- mass of NH_3 = ? g NH_3
- molar mass of NH_3 = 17.04 g/mol
- From the balanced equation: 3 mol H_2 = 2 mol NH_3.

2 Plan your work.

- To change grams of H_2 to moles, use the molar mass of H_2.
- The mole ratio must cancel out the units of mol H_2 given in the problem and leave the units of mol NH_3. Therefore, the mole ratio is

$$\frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}$$

- To change moles of NH_3 to grams, use the molar mass of NH_3.

3 Calculate.

$$? \text{ g } NH_3 = 1221 \cancel{\text{ g } H_2} \times \frac{1 \cancel{\text{ mol } H_2}}{2.02 \cancel{\text{ g } H_2}} \times \frac{2 \cancel{\text{ mol } NH_3}}{3 \cancel{\text{ mol } H_2}} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{ mol } NH_3}} = 6867 \text{ g } NH_3$$

4 Verify your result.

The units cancel to give the correct units for the answer. Estimating shows the answer should be about 6 times the original mass.

PRACTICE HINT

Remember to check both the units and the substance when canceling. For example, 1221 g H_2 cannot be converted to moles by multiplying by 1 mol NH_3/17.04 g NH_3. The units of grams in each one cannot cancel because they involve different substances.

PRACTICE

Use the equation below to answer the questions that follow.

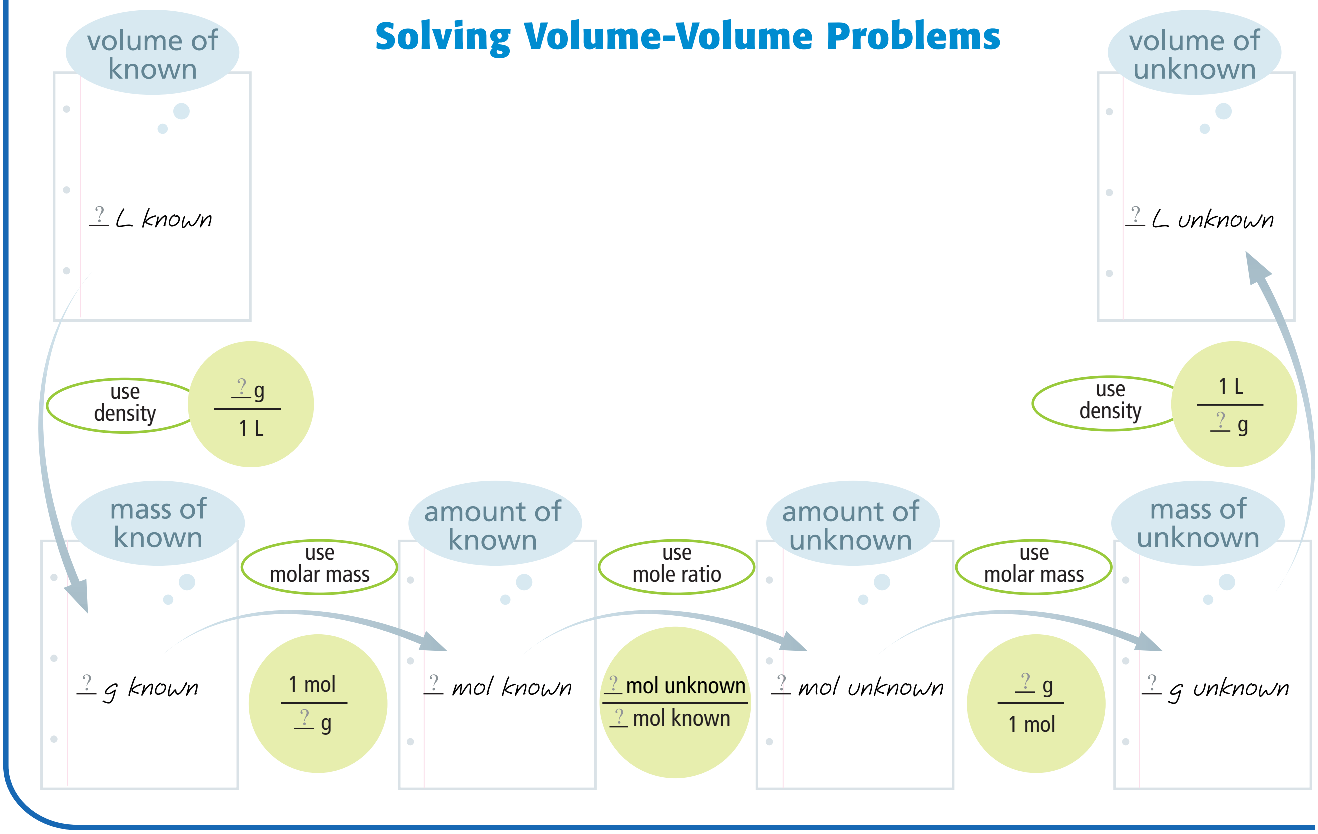
$$Fe_2O_3 + 2Al \longrightarrow 2Fe + Al_2O_3$$

1. How many grams of Al are needed to completely react with 135 g Fe_2O_3?
2. How many grams of Al_2O_3 can form when 23.6 g Al react with excess Fe_2O_3?
3. How many grams of Fe_2O_3 react with excess Al to make 475 g Fe?
4. How many grams of Fe will form when 97.6 g Al_2O_3 form?

PROBLEM SOLVING SKILL

4 SKILLS Toolkit

Solving Volume-Volume Problems



For Volume, You Might Use Density and Molar Mass

When reactants are liquids, they are almost always measured by volume. So, to do calculations involving liquids, you add two more steps to the sequence of mass-mass problems—the conversions of volume to mass and of mass to volume. Five conversion factors—two densities, two molar masses, and a mole ratio—are needed for this type of calculation, as shown in **Skills Toolkit 4.**

To convert from volume to mass or from mass to volume of a substance, use the density of the substance as the conversion factor. Keep in mind that the units you want to cancel should be on the bottom of your conversion factor.

There are ways other than density to include volume in stoichiometry problems. For example, if a substance in the problem is a gas at standard temperature and pressure (STP), use the molar volume of a gas to change directly between volume of the gas and moles. The molar volume of a gas is 22.41 L/mol for any gas at STP. Also, if a substance in the problem is in aqueous solution, then use the concentration of the solution to convert the volume of the solution to the moles of the substance dissolved. This procedure is especially useful when you perform calculations involving the reaction between an acid and a base. Of course, even in these problems, the basic process remains the same: change to moles, use the mole ratio, and change to the desired units.

SAMPLE PROBLEM C

Problems Involving Volume

What volume of H_3PO_4 forms when 56 mL $POCl_3$ completely react? (density of $POCl_3$ = 1.67 g/mL; density of H_3PO_4 = 1.83 g/mL)

$$POCl_3(l) + 3H_2O(l) \longrightarrow H_3PO_4(l) + 3HCl(g)$$

1 Gather information.

- volume $POCl_3$ = 56 mL $POCl_3$
- density $POCl_3$ = 1.67 g/mL
- molar mass $POCl_3$ = 153.32 g/mol
- volume H_3PO_4 = ?
- density H_3PO_4 = 1.83 g/mL
- molar mass H_3PO_4 = 98.00 g/mol
- From the equation: 1 mol $POCl_3$ = 1 mol H_3PO_4.

2 Plan your work.

- To change milliliters of $POCl_3$ to moles, use the density of $POCl_3$ followed by its molar mass.
- The mole ratio must cancel out the units of mol $POCl_3$ given in the problem and leave the units of mol H_3PO_4. Therefore, the mole ratio is

$$\frac{1 \text{ mol } H_3PO_4}{1 \text{ mol } POCl_3}$$

- To change out of moles of H_3PO_4 into milliliters, use the molar mass of H_3PO_4 followed by its density.

3 Calculate.

$$? \text{ mL } H_3PO_4 = 56 \cancel{\text{ mL } POCl_3} \times \frac{1.67 \cancel{\text{ g } POCl_3}}{1 \cancel{\text{ mL } POCl_3}} \times \frac{1 \cancel{\text{ mol } POCl_3}}{153.32 \cancel{\text{ g } POCl_3}} \times$$

$$\frac{1 \cancel{\text{ mol } H_3PO_4}}{1 \cancel{\text{ mol } POCl_3}} \times \frac{98.00 \cancel{\text{ g } H_3PO_4}}{1 \cancel{\text{ mol } H_3PO_4}} \times \frac{1 \text{ mL } H_3PO_4}{1.83 \cancel{\text{ g } H_3PO_4}} = 33 \text{ mL } H_3PO_4$$

4 Verify your result.

The units of the answer are correct. Estimating shows the answer should be about two-thirds of the original volume.

PRACTICE HINT

Do not try to memorize the exact steps of every type of problem. For long problems like these, you might find it easier to break the problem into three steps rather than solving all at once. Remember that whatever you are given, you need to change to moles, then use the mole ratio, then change out of moles to the desired units.

PRACTICE

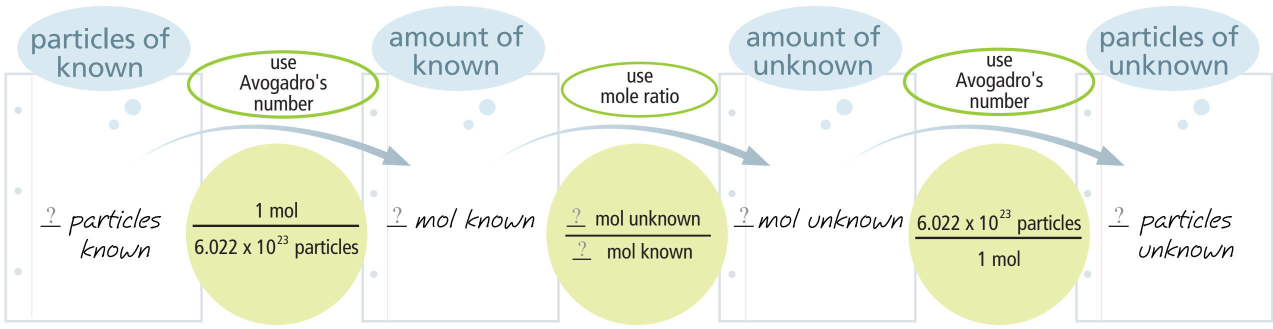
Use the densities and balanced equation provided to answer the questions that follow. (density of C_5H_{12} = 0.620 g/mL; density of C_5H_8 = 0.681 g/mL; density of H_2 = 0.0899 g/L)

$$C_5H_{12}(l) \longrightarrow C_5H_8(l) + 2H_2(g)$$

1. How many milliliters of C_5H_8 can be made from 366 mL C_5H_{12}?
2. How many liters of H_2 can form when 4.53×10^3 mL C_5H_8 form?
3. How many milliliters of C_5H_{12} are needed to make 97.3 mL C_5H_8?
4. How many milliliters of H_2 can be made from 1.98×10^3 mL C_5H_{12}?

PROBLEM SOLVING SKILL

Solving Particle Problems



Topic Link

Refer to the chapter "The Mole and Chemical Composition" for more information about Avogadro's number and molar mass.

For Number of Particles, Use Avogadro's Number

Skills Toolkit 5 shows how to use Avogadro's number, 6.022×10^{23} particles/mol, in stoichiometry problems. If you are given particles and asked to find particles, Avogadro's number cancels out! For this calculation you use only the coefficients from the balanced equation. In effect, you are interpreting the equation in terms of the number of particles again.

SAMPLE PROBLEM D

Problems Involving Particles

How many grams of C_5H_8 form from 1.89×10^{24} molecules C_5H_{12}?

$$C_5H_{12}(l) \longrightarrow C_5H_8(l) + 2H_2(g)$$

PRACTICE HINT

Expect more problems like this one that do not exactly follow any single **Skills Toolkit** in this chapter. These problems will combine steps from one or more problems, but all will still use the mole ratio as the key step.

1 Gather information.

- quantity of $C_5H_{12} = 1.89 \times 10^{24}$ molecules
- Avogadro's number $= 6.022 \times 10^{23}$ molecules/mol
- mass of $C_5H_8 = ?$ g C_5H_8
- molar mass of $C_5H_8 = 68.13$ g/mol
- From the balanced equation: 1 mol C_5H_{12} = 1 mol C_5H_8.

2 Plan your work.

Set up the problem using Avogadro's number to change to moles, then use the mole ratio, and finally use the molar mass of C_5H_8 to change to grams.

3 Calculate.

$$? \text{ g } C_5H_8 = 1.89 \times 10^{24} \cancel{\text{molecules } C_5H_{12}} \times \frac{1 \cancel{\text{ mol } C_5H_{12}}}{6.022 \times 10^{23} \cancel{\text{ molecules } C_5H_{12}}} \times$$

$$\frac{1 \cancel{\text{ mol } C_5H_8}}{1 \cancel{\text{ mol } C_5H_{12}}} \times \frac{68.13 \text{ g } C_5H_8}{1 \cancel{\text{ mol } C_5H_8}} = 214 \text{ g } C_5H_8$$

4 Verify your result.

The units cancel correctly, and estimating gives 210.

PRACTICE

PROBLEM SOLVING SKILL

Use the equation provided to answer the questions that follow.

$$Br_2(l) + 5F_2(g) \longrightarrow 2BrF_5(l)$$

1 How many molecules of BrF_5 form when 384 g Br_2 react with excess F_2?

2 How many molecules of Br_2 react with 1.11×10^{20} molecules F_2?

Many Problems, Just One Solution

Although you could be given many different problems, the solution boils down to just three steps. Take whatever you are given, and find a way to change it into moles. Then, use a mole ratio from the balanced equation to get moles of the second substance. Finally, find a way to convert the moles into the units that you need for your final answer.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What conversion factor is present in almost all stoichiometry calculations?

2. For a given substance, what information links mass to moles? number of particles to moles?

3. What conversion factor will change moles CO_2 to grams CO_2? moles H_2O to molecules H_2O?

PRACTICE PROBLEMS

4. Use the equation below to answer the questions that follow.

$$Br_2 + Cl_2 \longrightarrow 2BrCl$$

a. How many moles of BrCl form when 2.74 mol Cl_2 react with excess Br_2?

b. How many grams of BrCl form when 239.7 g Cl_2 react with excess Br_2?

c. How many grams of Br_2 are needed to react with 4.53×10^{25} molecules Cl_2?

5. The equation for burning C_2H_2 is

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

a. If 15.9 L C_2H_2 react at STP, how many moles of CO_2 are produced? (Hint: At STP, 1 mol = 22.41 L for any gas.)

b. How many milliliters of CO_2 (density = 1.977 g/L) can be made when 59.3 mL O_2 (density = 1.429 g/L) react?

CRITICAL THINKING

6. Why do you need to use amount in moles to solve stoichiometry problems? Why can't you just convert from mass to mass?

7. LiOH and NaOH can each react with CO_2 to form the metal carbonate and H_2O. These reactions can be used to remove CO_2 from the air in a spacecraft.

a. Write a balanced equation for each reaction.

b. Calculate the grams of NaOH and of LiOH that remove 288 g CO_2 from the air.

c. NaOH is less expensive per mole than LiOH. Based on your calculations, explain why LiOH is used during shuttle missions rather than NaOH.

SECTION 2

Limiting Reactants and Percentage Yield

KEY TERMS

- **limiting reactant**
- **excess reactant**
- **actual yield**

OBJECTIVES

1. **Identify** the limiting reactant for a reaction and use it to calculate theoretical yield.
2. **Perform** calculations involving percentage yield.

Limiting Reactants and Theoretical Yield

To drive a car, you need gasoline in the tank and oxygen from the air. When the gasoline runs out, you can't go any farther even though there is still plenty of oxygen. In other words, the gasoline limits the distance you can travel because it runs out and the reaction in the engine stops.

In the previous section, you assumed that 100% of the reactants changed into products. And that is what should happen theoretically. But in the real world, other factors, such as the amounts of all reactants, the completeness of the reaction, and product lost in the process, can limit the yield of a reaction. The analogy of assembling homecoming mums for a fund raiser, as shown in **Figure 3,** will help you understand that whatever is in short supply will limit the quantity of product made.

Figure 3
The number of mums these students can assemble will be limited by the component that runs out first.

The Limiting Reactant Forms the Least Product

The students assembling mums use one helmet, one flower, eight blue ribbons, six white ribbons, and two bells to make each mum. As a result, the students cannot make any more mums once any one of these items is used up. Likewise, the reactants of a reaction are seldom present in ratios equal to the mole ratio in the balanced equation. So one of the reactants is used up first. For example, one way to to make H_2 is

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

If you combine 0.23 mol Zn and 0.60 mol HCl, would they react completely? Using the coefficients from the balanced equation, you can predict that 0.23 mol Zn can form 0.23 mol H_2, and 0.60 mol HCl can form 0.30 mol H_2. Zinc is called the **limiting reactant** because the zinc limits the amount of product that can form. The zinc is used up first by the reaction. The HCl is the **excess reactant** because there is more than enough HCl present to react with all of the Zn. There will be some HCl left over after the reaction stops.

Again, think of the mums, and look at **Figure 4.** The supplies at left are the available reactants. The products formed are the finished mums. The limiting reactant is the flowers because they are completely used up first. The ribbons, helmets, and bells are excess reactants because there are some of each of these items left over, at right. You can determine the limiting reactant by calculating the amount of product that each reactant could form. Whichever reactant would produce the least amount of product is the limiting reactant.

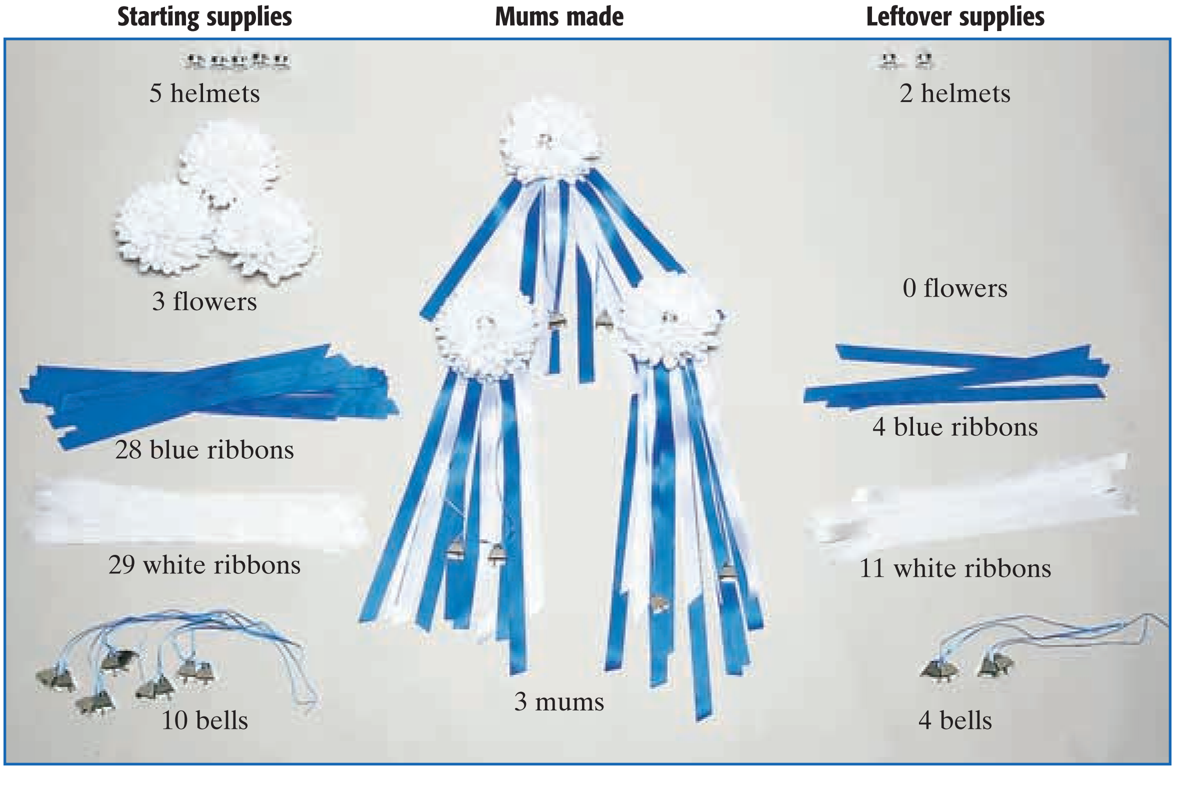
limiting reactant

the substance that controls the quantity of product that can form in a chemical reaction

excess reactant

the substance that is not used up completely in a reaction

Figure 4
The flowers are in short supply. They are the limiting reactant for assembling these homecoming mums.



Determine Theoretical Yield from the Limiting Reactant

So far you have done only calculations that assume reactions happen perfectly. The maximum quantity of product that a reaction could theoretically make if everything about the reaction works perfectly is called the *theoretical yield*. The theoretical yield of a reaction should always be calculated based on the limiting reactant.

In the reaction of Zn with HCl, the theoretical yield is 0.23 mol H_2 even though the HCl could make 0.30 mol H_2.

SAMPLE PROBLEM E

Limiting Reactants and Theoretical Yield

Identify the limiting reactant and the theoretical yield of phosphorous acid, H_3PO_3, if 225 g of PCl_3 is mixed with 125 g of H_2O.

$$PCl_3 + 3H_2O \longrightarrow H_3PO_3 + 3HCl$$

1 Gather information.

- mass PCl_3 = 225 g PCl_3
- mass H_2O = 125 g H_2O
- mass H_3PO_3 = ? g H_3PO_3
- From the balanced equation:
  1 mol PCl_3 = 1 mol H_3PO_3 and 3 mol H_2O = 1 mol H_3PO_3.
- molar mass PCl_3 = 137.32 g/mol
- molar mass H_2O = 18.02 g/mol
- molar mass H_3PO_3 = 82.00 g/mol

2 Plan your work.

Set up problems that will calculate the mass of H_3PO_3 you would expect to form from each reactant.

3 Calculate.

$$? \text{ g } H_3PO_3 = 225 \cancel{\text{ g } PCl_3} \times \frac{1 \cancel{\text{ mol } PCl_3}}{137.32 \cancel{\text{ g } PCl_3}} \times \frac{1 \cancel{\text{ mol } H_3PO_3}}{1 \cancel{\text{ mol } PCl_3}} \times \frac{82.00 \text{ g } H_3PO_3}{1 \cancel{\text{ mol } H_3PO_3}} = 134 \text{ g } H_3PO_3$$

$$? \text{ g } H_3PO_3 = 123 \cancel{\text{ g } H_2O} \times \frac{1 \cancel{\text{ mol } H_2O}}{18.02 \cancel{\text{ g } H_2O}} \times \frac{1 \cancel{\text{ mol } H_3PO_3}}{3 \cancel{\text{ mol } H_2O}} \times \frac{82.00 \text{ g } H_3PO_3}{1 \cancel{\text{ mol } H_3PO_3}} = 187 \text{ g } H_3PO_3$$

PCl_3 is the limiting reactant. The theoretical yield is 134 g H_3PO_3.

4 Verify your result.

The units of the answer are correct, and estimating gives 128.

PRACTICE HINT

Whenever a problem gives you quantities of two or more reactants, you must determine the limiting reactant and use it to determine the theoretical yield.

PRACTICE

Using the reaction above, identify the limiting reactant and the theoretical yield (in grams) of HCl for each pair of reactants.

1. 3.00 mol PCl_3 and 3.00 mol H_2O
2. 75.0 g PCl_3 and 75.0 g H_2O
3. 1.00 mol of PCl_3 and 50.0 g of H_2O

PROBLEM SOLVING SKILL

Limiting Reactants and the Food You Eat

In industry, the cheapest reactant is often used as the excess reactant. In this way, the expensive reactant is more completely used up. In addition to being cost-effective, this practice can be used to control which reactions happen. In the production of cider vinegar from apple juice, the apple juice is first kept where there is no oxygen so that the microorganisms in the juice break down the sugar, glucose, into ethanol and carbon dioxide. The resulting solution is hard cider.

Having excess oxygen in the next step allows the organisms to change ethanol into acetic acid, resulting in cider vinegar. Because the oxygen in the air is free and is easy to get, the makers of cider vinegar constantly pump air through hard cider as they make it into vinegar. Ethanol, which is not free, is the limiting reactant and is used up in the reaction.

Cost is also used to choose the excess reactant when making banana flavoring, isopentyl acetate. Acetic acid is the excess reactant because it costs much less than isopentyl alcohol.

$$CH_3COOH + C_5H_{11}OH \longrightarrow CH_3COOC_5H_{11} + H_2O$$
$$\text{acetic acid} + \text{isopentyl alcohol} \longrightarrow \text{isopentyl acetate} + \text{water}$$

As shown in **Figure 5,** when compared mole for mole, isopentyl alcohol is more than twice as expensive as acetic acid. When a large excess of acetic acid is present, almost all of the isopentyl alcohol reacts.

Choosing the excess and limiting reactants based on cost is also helpful in areas outside of chemistry. In making the homecoming mums, the flower itself is more expensive than any of the other materials, so it makes sense to have an excess of ribbons and charms. The expensive flowers are the limiting reactant.

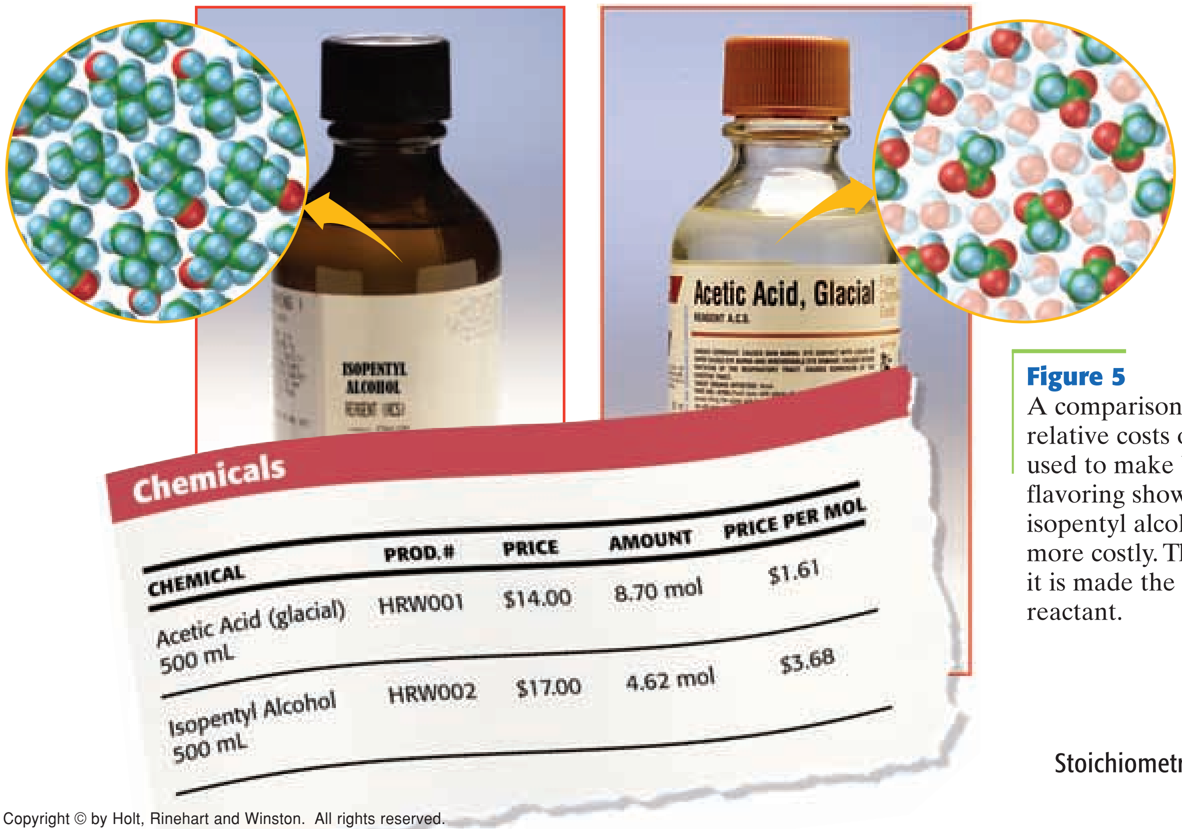


Figure 5
A comparison of the relative costs of chemicals used to make banana flavoring shows that isopentyl alcohol is more costly. That is why it is made the limiting reactant.

Table 1 Predictions and Results for Isopentyl Acetate Synthesis

Reactants	Formula	Mass present	Amount present	Amount left over
Isopentyl alcohol	$C_5H_{11}OH$	500.0 g	5.67 mol (limiting reactant)	0.0 mol
Acetic acid	CH_3COOH	1.25×10^3 g	20.8 mol	15.1 mol
Products	**Formula**	**Amount expected**	**Theoretical yield (mass expected)**	**Actual yield (mass produced)**
Isopentyl acetate	$CH_3COOC_5H_{11}$	5.67 mol	738 g	591 g
Water	H_2O	5.67 mol	102 g	81.6 g

Actual Yield and Percentage Yield

actual yield

the measured amount of a product of a reaction

Although equations tell you what *should* happen in a reaction, they cannot always tell you what *will* happen. For example, sometimes reactions do not make all of the product predicted by stoichiometric calculations, or the theoretical yield. In most cases, the **actual yield,** the mass of product actually formed, is less than expected. Imagine that a worker at the flavoring factory mixes 500.0 g isopentyl alcohol with 1.25×10^3 g acetic acid. The actual and theoretical yields are summarized in **Table 1.** Notice that the actual yield is less than the mass that was expected.

There are several reasons why the actual yield is usually less than the theoretical yield in chemical reactions. Many reactions do not completely use up the limiting reactant. Instead, some of the products turn back into reactants so that the final result is a mixture of reactants and products. In many cases the main product must go through additional steps to purify or separate it from other chemicals. For example, banana flavoring must be *distilled,* or isolated based on its boiling point. Solid compounds, such as sugar, must be recrystallized. Some of the product may be lost in the process. There also may be other reactions, called *side reactions*, that can use up reactants without making the desired product.

Determining Percentage Yield

The ratio relating the actual yield of a reaction to its theoretical yield is called the *percentage yield* and describes the efficiency of a reaction. Calculating a percentage yield is similar to calculating a batting average. A batter might get a hit every time he or she is at bat. This is the "theoretical yield." But no player has gotten a hit every time. Suppose a batter gets 41 hits in 126 times at bat. The batting average is 41 (the actual hits) divided by 126 (the possible hits theoretically), or 0.325. In the example described in **Table 1,** the theoretical yield for the reaction is 738 g. The actual yield is 591 g. The percentage yield is

$$\text{percentage yield} = \frac{591 \text{ g (actual yield)}}{738 \text{ g (theoretical yield)}} \times 100 = 80.1\%$$

SAMPLE PROBLEM F

Calculating Percentage Yield

Determine the limiting reactant, the theoretical yield, and the percentage yield if 14.0 g N_2 are mixed with 9.0 g H_2, and 16.1 g NH_3 form.

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

1 Gather information.

- mass N_2 = 14.0 g N_2
- mass H_2 = 9.0 g H_2
- theoretical yield of NH_3 = ? g NH_3
- actual yield of NH_3 = 16.1 g NH_3
- From the balanced equation:
  1 mol N_2 = 2 mol NH_3 and 3 mol H_2 = 2 mol NH_3.
- molar mass N_2 = 28.02 g/mol
- molar mass H_2 = 2.02 g/mol
- molar mass NH_3 = 17.04 g/mol

2 Plan your work.

Set up problems that will calculate the mass of NH_3 you would expect to form from each reactant.

3 Calculate.

$$? \text{ g } NH_3 = 14.0 \text{ g } \cancel{N_2} \times \frac{1 \cancel{\text{ mol } N_2}}{28.02 \text{ g } \cancel{N_2}} \times \frac{2 \cancel{\text{ mol } NH_3}}{1 \cancel{\text{ mol } N_2}} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{ mol } NH_3}} = 17.0 \text{ g } NH_3$$

$$? \text{ g } NH_3 = 9.0 \text{ g } \cancel{H_2} \times \frac{1 \cancel{\text{ mol } H_2}}{2.02 \text{ g } \cancel{H_2}} \times \frac{2 \cancel{\text{ mol } NH_3}}{3 \cancel{\text{ mol } H_2}} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{ mol } NH_3}} = 51 \text{ g } NH_3$$

- The smaller quantity made, 17.0 g NH_3, is the theoretical yield so the limiting reactant is N_2.
- The percentage yield is calculated:

$$\text{percentage yield} = \frac{16.1 \text{ g (actual yield)}}{17.0 \text{ g (theoretical yield)}} \times 100 = 94.7\%$$

4 Verify your result.

The units of the answer are correct. The percentage yield is less than 100%, so the final calculation is probably set up correctly.

PRACTICE HINT

If an amount of product actually formed is given in a problem, this is the reaction's actual yield.

PRACTICE

Determine the limiting reactant and the percentage yield for each of the following.

1. 14.0 g N_2 react with 3.15 g H_2 to give an actual yield of 14.5 g NH_3.
2. In a reaction to make ethyl acetate, 25.5 g CH_3COOH react with 11.5 g C_2H_5OH to give a yield of 17.6 g $CH_3COOC_2H_5$.

$$CH_3COOH + C_2H_5OH \longrightarrow CH_3COOC_2H_5 + H_2O$$

3. 16.1 g of bromine are mixed with 8.42 g of chlorine to give an actual yield of 21.1 g of bromine monochloride.

PROBLEM SOLVING SKILL

Determining Actual Yield

Although the actual yield can only be determined experimentally, a close estimate can be calculated if the percentage yield for a reaction is known. The percentage yield in a particular reaction is usually fairly consistent. For example, suppose an industrial chemist determined the percentage yield for six tries at making banana flavoring and found the results were 80.0%, 82.1%, 79.5%, 78.8%, 80.5%, and 81.9%. In the future, the chemist can expect a yield of around 80.5%, or the average of these results.

If the chemist has enough isopentyl alcohol to make 594 g of the banana flavoring theoretically, then an actual yield of around 80.5% of that, or 478 g, can be expected.

SAMPLE PROBLEM G

Calculating Actual Yield

How many grams of $CH_3COOC_5H_{11}$ should form if 4808 g are theoretically possible and the percentage yield for the reaction is 80.5%?

1 Gather information.

- theoretical yield of $CH_3COOC_5H_{11}$ = 4808 g $CH_3COOC_5H_{11}$
- actual yield of $CH_3COOC_5H_{11}$ = ? g $CH_3COOC_5H_{11}$
- percentage yield = 80.5%

2 Plan your work.

Use the percentage yield and the theoretical yield to calculate the actual yield expected.

3 Calculate.

$$80.5\% = \frac{\text{actual yield}}{4808 \text{ g}} \times 100$$

actual yield = 4808 g × 0.805 = 3.87×10^3 g $CH_3COOC_5H_{11}$

4 Verify your result.

The units of the answer are correct. The actual yield is less than the theoretical yield, as it should be.

PRACTICE HINT

The actual yield should always be less than the theoretical yield. A wrong answer that is greater than the theoretical yield can result if you accidentally reverse the actual and theoretical yields.

PRACTICE

PROBLEM SOLVING SKILL

1. The percentage yield of NH_3 from the following reaction is 85.0%. What actual yield is expected from the reaction of 1.00 kg N_2 with 225 g H_2?

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

2. If the percentage yield is 92.0%, how many grams of CH_3OH can be made by the reaction of 5.6×10^3 g CO with 1.0×10^3 g H_2?

$$CO + 2H_2 \longrightarrow CH_3OH$$

3. Suppose that the percentage yield of BrCl is 90.0%. How much BrCl can be made by reacting 338 g Br_2 with 177 g Cl_2?