12
Stoichiometry

INSIDE:

- **12.1** The Arithmetic of Equations
- **12.2** Chemical Calculations
- **12.3** Limiting Reagent and Percent Yield

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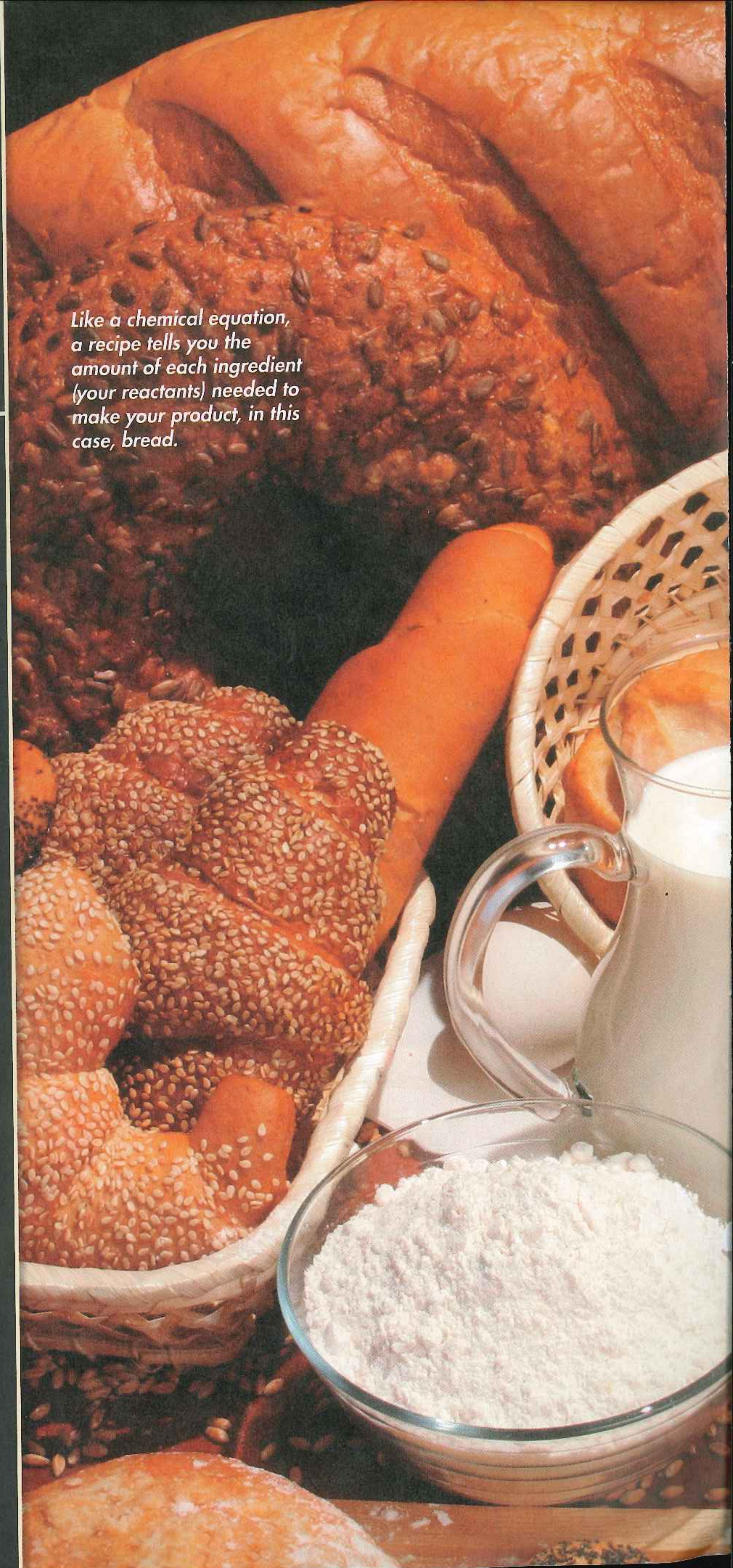
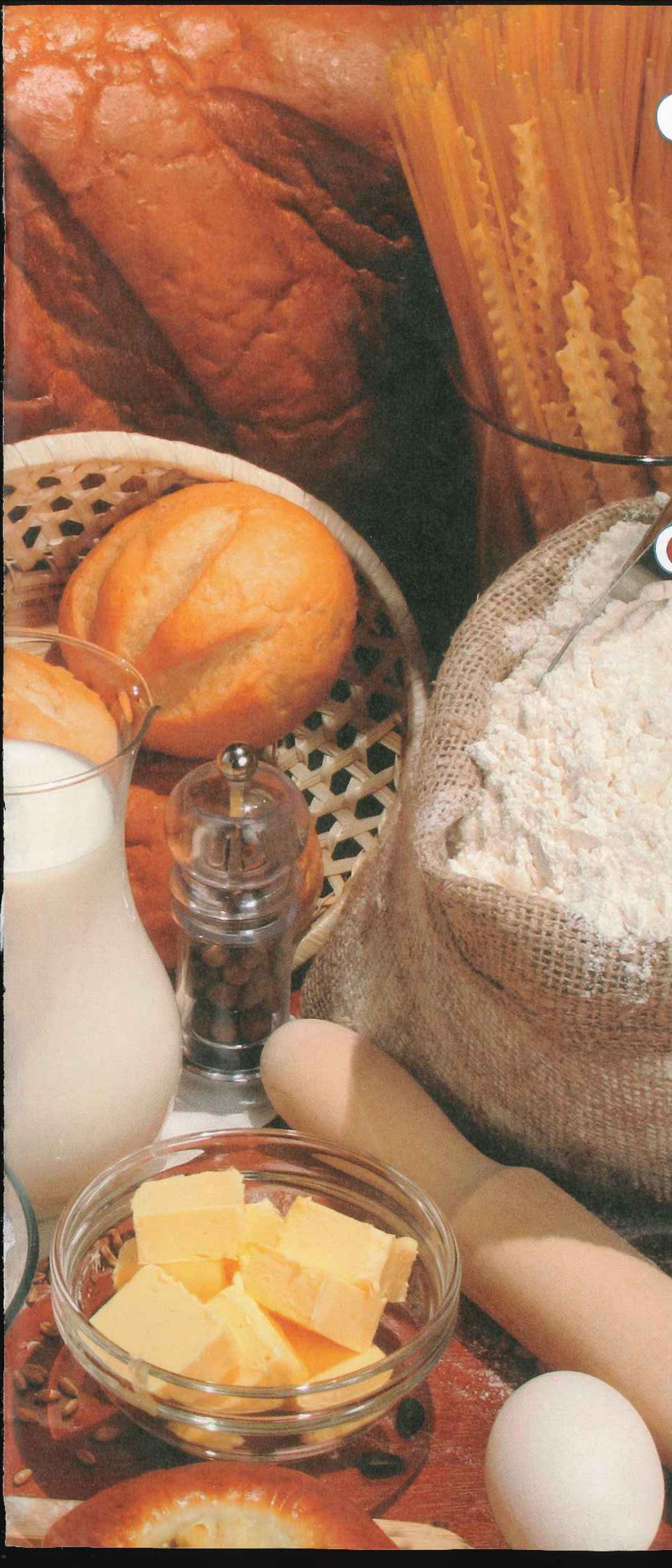


Like a chemical equation, a recipe tells you the amount of each ingredient (your reactants) needed to make your product, in this case, bread.



BIGIDEAS

- **THE MOLE**
- **REACTIONS**

Essential Questions:

1. *How are balanced chemical equations used in stoichiometric calculations?*
2. *How can you calculate amounts of reactants and products in a chemical reaction?*

CHEMYSTERY

Cookie Crumbles

For the school bake sale, Jack wanted to make cookies to sell. He looked in cookbooks to find a good recipe. The recipe he chose called for specific amounts of butter, flour, sugar, eggs, vanilla, and baking soda. Jack wanted to make sure that his cookies were delicious and sweet. He didn't think there was enough sugar in the recipe, so he added twice as much sugar as the recipe called for.

Jack mixed the ingredients, put balls of the dough on a cookie sheet, and placed them in the oven to bake. When the bake time was up, Jack was very disappointed in his cookies. Instead of sweet, delicious cookies, his cookies were brown, hard, and crumbly. What happened? He checked the oven temperature and the amount of time that the cookies were in the oven. The time and temperature matched the directions in the recipe. Why didn't Jack's cookies turn out as he expected?

► Connect to the **BIG**IDEA As you read about quantifying chemical reactions, think about what could have happened to Jack's cookies.

NATIONAL SCIENCE EDUCATION STANDARDS

B-3

12.1 The Arithmetic of Equations

CHEMISTRY & YOU

Q: *How do you figure out how much starting material you need to make a finished product?* Whenever you make something, you need to have the ingredients or the parts that make up the desired product. When making bikes, you need parts such as wheels, handlebars, pedals, and frames. If a factory needs to make 200 bikes, then the workers would need to calculate how many of each part they need to produce the 200 bikes. In this lesson, you will learn about how chemists determine how much of each reactant is needed to make a certain amount of product.

Key Questions

How do chemists use balanced chemical equations?

In terms of what quantities can you interpret a balanced chemical equation?

Vocabulary

- stoichiometry

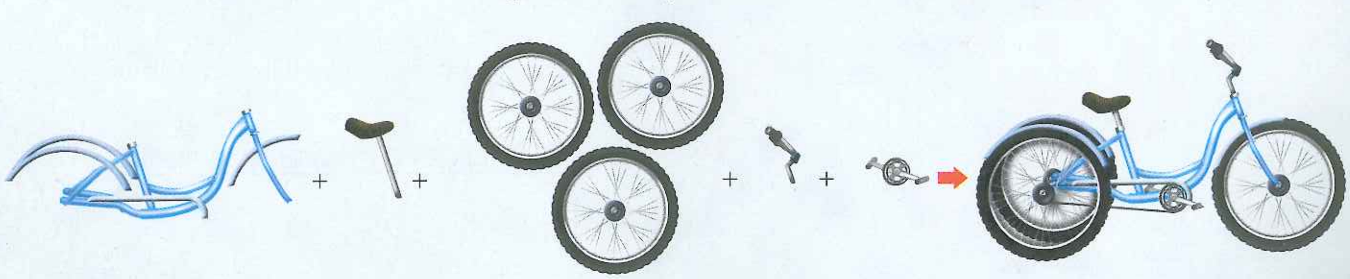
Using Equations

How do chemists use balanced chemical equations?

One example of something that you might make is food. When you make cookies, for instance, you probably use a recipe. A cookie recipe tells you the precise amounts of ingredients to mix to make a certain number of cookies. If you need a larger number of cookies than the recipe provides for, you can double or triple the amounts of all the ingredients. In a way, a cookie recipe provides the same kind of information that a balanced chemical equation provides. In a cookie recipe, you can think of the ingredients as the reactants and the cookies as the products.

Everyday Equations The making of tricycles, like bikes and cookies, is a job that requires quantitative information to create the final product. Let's say you are in charge of manufacturing for the Travel Time Tricycle Company. The business plan for Travel Time requires the production of 640 custom-made tricycles each week. One of your responsibilities is to make sure there are enough parts available at the start of each workweek to make these tricycles. How can you determine the number of parts you need per week?

To simplify this discussion, assume that the major components of the tricycle are the frame (F), the seat (S), the wheels (W), the handlebars (H), and the pedals (P)—in other words, the reactants. The figure below illustrates how an equation can represent the manufacturing of a single tricycle.



$$F + S + 3W + H + 2P \longrightarrow FSW_3HP_2$$

The finished tricycle, your product, has a "formula" of FSW_3HP_2. The balanced equation for making a single tricycle is

$$F + S + 3W + H + 2P \longrightarrow FSW_3HP_2$$

This balanced equation is a "recipe" to make a single tricycle: Making a tricycle requires assembling one frame, one seat, three wheels, one handlebar, and two pedals. Now look at Sample Problem 12.1. It shows you how to use the balanced equation to calculate the number of parts needed to manufacture a given number of tricycles.

CHEM TUTOR

SampleProblem 12.1

Using a Balanced Equation as a Recipe

In a five-day workweek, Travel Time is scheduled to make 640 tricycles. How many wheels should be in the plant on Monday morning to make these tricycles?

1 Analyze **List the knowns and the unknown.** Use the balanced equation to identify a conversion factor that will allow you to calculate the unknown. The conversion you need to make is from tricycles (FSW_3HP_2) to wheels (W).

KNOWNS

number of tricycles = 640 tricycles = 640 FSW_3HP_2

$F + S + 3W + H + 2P \longrightarrow FSW_3HP_2$

UNKNOWN

number of wheels = ? W

2 Calculate **Solve for the unknown.**

Identify a conversion factor that relates wheels to tricycles. You can write two conversion factors relating wheels to tricycles.

$$\frac{3\ W}{1\ FSW_3HP_2} \text{ and } \frac{1\ FSW_3HP_2}{3\ W}$$

The desired unit is W; so use the conversion factor on the left. Multiply the number of tricycles by the conversion factor.

$$640\ \cancel{FSW_3HP_2} \times \frac{3\ W}{1\ \cancel{FSW_3HP_2}} = \boxed{1920\ W}$$

When using conversion factors, remember to cancel like units when they are in both the numerator and denominator. This tells you that you are using the correct conversion factor.

3 Evaluate **Does the result make sense?** If three wheels are required for each tricycle and more than 600 tricycles are being made, then a number of wheels in excess of 1800 is a logical answer. The unit of the known (FSW_3HP_2) cancels, and the answer has the correct unit (W).

1. Travel Time has decided to make 288 tricycles each day. How many tricycle seats, wheels, and pedals are needed for each day?

2. Write an equation that gives your own "recipe" for making a skateboard.

CHEMISTRY & YOU

Q: *How can you determine the amount of each reactant you need to make a product?*

CONCEPTS IN ACTION *Learn more about* **stoichiometry** *online.*

Balanced Chemical Equations Nearly everything you use is manufactured from chemicals—soaps, shampoos and conditioners, CDs, cosmetics, medicines, and clothes. When manufacturing such items, the cost of making them cannot be greater than the price at which they are sold. Otherwise, the manufacturer will not make a profit. Therefore, the chemical processes used in manufacturing must be carried out economically. A situation like this is where balanced equations help.

A balanced chemical equation tells you what amounts of reactants to mix and what amount of a product to expect. **Chemists use balanced chemical equations as a basis to calculate how much reactant is needed or how much product will be formed in a reaction.** When you know the quantity of one substance in a reaction, you can calculate the quantity of any other substance consumed or created in the reaction. Quantity usually means the amount of a substance expressed in grams or moles. However, quantity could just as well be in liters, tons, or molecules.

The calculation of quantities in chemical reactions is a subject of chemistry called **stoichiometry.** Calculations using balanced equations are called stoichiometric calculations. For chemists, stoichiometry is a form of bookkeeping. For example, accountants can track income, expenditures, and profits for a small business by tallying each in dollars and cents. Chemists can track reactants and products in a reaction by stoichiometry. It allows chemists to tally the amounts of reactants and products using ratios of moles or representative particles derived from chemical equations.

READING SUPPORT

Build Vocabulary: ***Word Origins*** *Stoichiometry* comes from the combination of the Greek words *stoikheioin,* meaning "element," and *metron,* meaning "to measure." Stoichiometry is the calculation of amounts of substances involved in chemical reactions. ***What do you first need to know about a chemical reaction before doing stoichiometry calculations?***

Chemical Equations

In terms of what quantities can you interpret a balanced chemical equation?

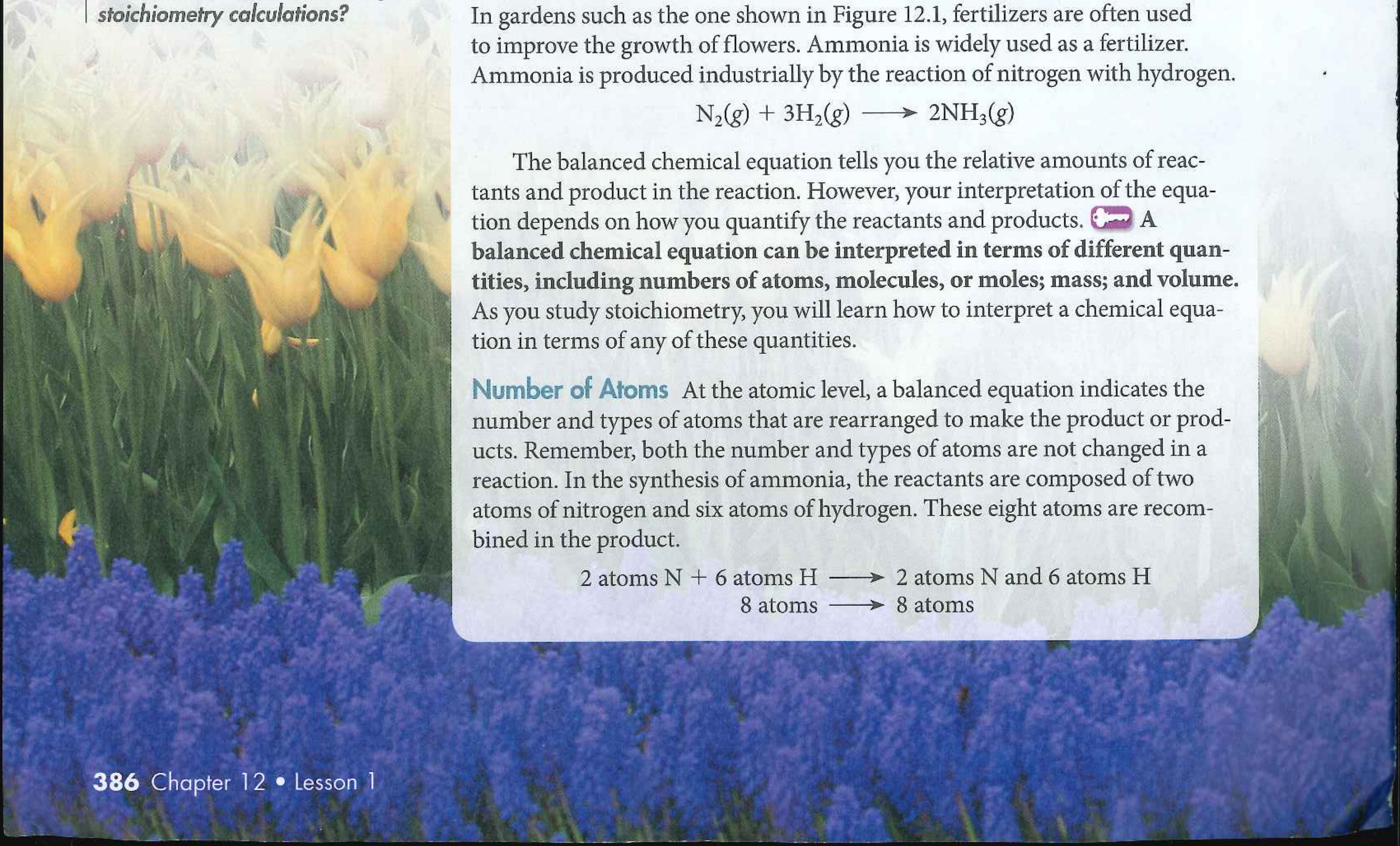
In gardens such as the one shown in Figure 12.1, fertilizers are often used to improve the growth of flowers. Ammonia is widely used as a fertilizer. Ammonia is produced industrially by the reaction of nitrogen with hydrogen.

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

The balanced chemical equation tells you the relative amounts of reactants and product in the reaction. However, your interpretation of the equation depends on how you quantify the reactants and products. **A balanced chemical equation can be interpreted in terms of different quantities, including numbers of atoms, molecules, or moles; mass; and volume.** As you study stoichiometry, you will learn how to interpret a chemical equation in terms of any of these quantities.

Number of Atoms At the atomic level, a balanced equation indicates the number and types of atoms that are rearranged to make the product or products. Remember, both the number and types of atoms are not changed in a reaction. In the synthesis of ammonia, the reactants are composed of two atoms of nitrogen and six atoms of hydrogen. These eight atoms are recombined in the product.

$$\begin{array}{rcl} 2 \text{ atoms N} + 6 \text{ atoms H} & \longrightarrow & 2 \text{ atoms N and 6 atoms H} \\ 8 \text{ atoms} & \longrightarrow & 8 \text{ atoms} \end{array}$$



Number of Molecules The balanced equation indicates that one molecule of nitrogen reacts with three molecules of hydrogen. Nitrogen and hydrogen will always react to form ammonia in a 1:3:2 ratio of molecules. If you could make 10 molecules of nitrogen react with 30 molecules of hydrogen, you would expect to get 20 molecules of ammonia. Of course, it is not practical to count such small numbers of molecules and allow them to react. You could, however, take Avogadro's number of nitrogen molecules and make them react with three times Avogadro's number of hydrogen molecules. This value would be the same 1:3 ratio of molecules of reactants. The reaction would form two times Avogadro's number of ammonia molecules.

$$1 \times \begin{array}{c}6.02 \times 10^{23}\\ \text{molecules } N_2\end{array} + 3 \times \begin{array}{c}6.02 \times 10^{23}\\ \text{molecules } H_2\end{array} \longrightarrow 2 \times \begin{array}{c}6.02 \times 10^{23}\\ \text{molecules } NH_3\end{array}$$

Moles You know that Avogrado's number of representative particles is equal to one mole of a substance. Therefore, since a balanced chemical equation tells you the number of representative particles, it also tells you the number of moles. The coefficients of a balanced chemical equation indicate the relative numbers of moles of reactants and products in a chemical reaction. These numbers are the most important pieces of information that a balanced chemical equation provides. Using this information, you can calculate the amounts of reactants and products. In the synthesis of ammonia, one mole of nitrogen molecules reacts with three moles of hydrogen molecules to form two moles of ammonia molecules. As you can see from this reaction, the total number of moles of reactants does not equal the total number of moles of product.

$$1 \text{ mol } N_2 + 3 \text{ mol } H_2 \longrightarrow 2 \text{ mol } NH_3$$

Mass A balanced chemical equation obeys the law of conservation of mass. This law states that mass can be neither created nor destroyed in an ordinary chemical or physical process. As you recall, the number and type of atoms does not change in a chemical reaction. Therefore, the total mass of the atoms in the reaction does not change. Using the mole relationship, you can relate mass to the number of atoms in the chemical equation. The mass of 1 mol of N_2 (28.0 g) plus the mass of 3 mol of H_2 (6.0 g) equals the mass of 2 mol of NH_3 (34.0 g). Although the number of moles of reactants does not equal the number of moles of product, the total number of grams of reactants does equal the total number of grams of product.

$$\begin{array}{c}28.0 \text{ g } N_2 + (3 \times 2.0 \text{ g } H_2) \longrightarrow (2 \times 17.0 \text{ g } NH_3)\\ 34.0 \text{ g} \longrightarrow 34.0 \text{ g}\end{array}$$

Volume If you assume standard temperature and pressure, the equation also tells you about the volumes of gases. Recall that 1 mol of any gas at STP occupies a volume of 22.4 L. The equation indicates that 22.4 L of N_2 reacts with 67.2 L (3 × 22.4 L) of H_2. This reaction forms 44.8 L (2 × 22.4 L) of NH_3.

$$22.4 \text{ L } N_2 + 67.2 \text{ L } H_2 \longrightarrow 44.8 \text{ L } NH_3$$

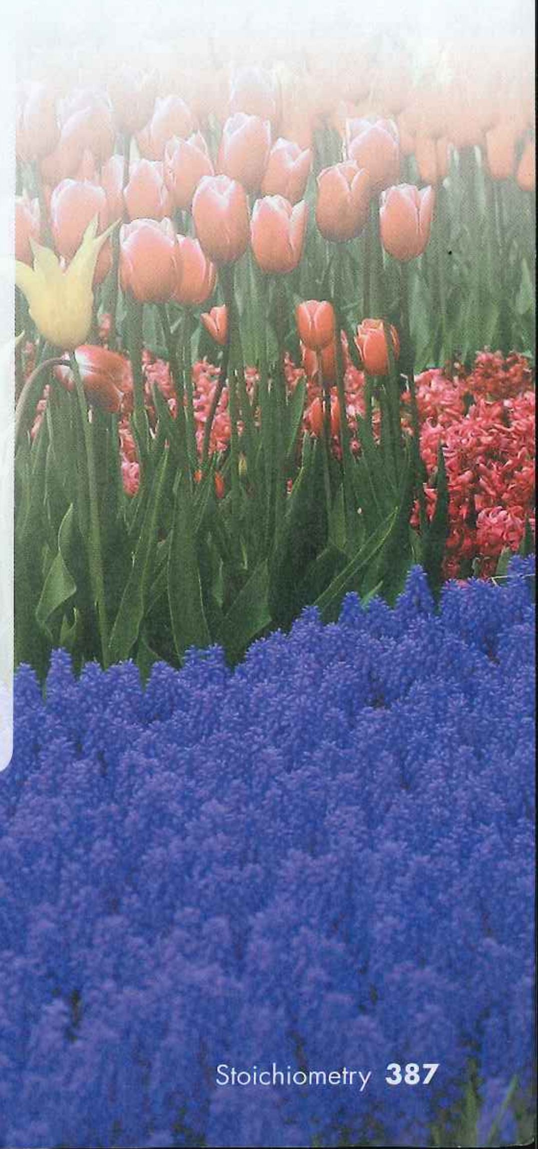


Figure 12.1 Use of Ammonia
Gardeners use ammonium salts as fertilizer. The nitrogen in these salts is essential to plant growth.

SampleProblem 12.2

Interpreting a Balanced Chemical Equation

Hydrogen sulfide, which smells like rotten eggs, is found in volcanic gases. The balanced equation for the burning of hydrogen sulfide is

$$2H_2S(g) + 3O_2(g) \longrightarrow 2SO_2(g) + 2H_2O(g)$$

Interpret this equation in terms of

a. numbers of representative particles and moles.

b. masses of reactants and products.

Remember that atoms and molecules are both representative particles. In this equation, all the reactants and products are molecules; so all the representative particles are molecules.

1 Analyze **Identify the relevant concepts.** The coefficients in the balanced equation give the relative number of representative particles and moles of reactants and products. A balanced chemical equation obeys the law of conservation of mass.

2 Solve **Apply concepts to this situation.**

Use the coefficients in the balanced equation to identify the number of representative particles and moles.

a. $2 \text{ molecules } H_2S + 3 \text{ molecules } O_2 \longrightarrow 2 \text{ molecules } SO_2 + 2 \text{ molecules } H_2O$

$2 \text{ mol } H_2S + 3 \text{ mol } O_2 \longrightarrow 2 \text{ mol } SO_2 + 2 \text{ mol } H_2O$

Use the periodic table to calculate the molar mass of each reactant and product.

b. $1 \text{ mol } H_2S = 34.1 \text{ g } H_2S$

$1 \text{ mol } O_2 = 32.0 \text{ g } O_2$

$1 \text{ mol } SO_2 = 64.1 \text{ g } SO_2$

$1 \text{ mol } H_2O = 18.0 \text{ g } H_2O$

Multiply the number of moles of each reactant and product by its molar mass.

$2 \text{ mol } H_2S + 3 \text{ mol } O_2 \longrightarrow 2 \text{ mol } SO_2 + 2 \text{ mol } H_2O.$

$$\left(2 \cancel{\text{mol}} \times 34.1 \frac{\text{g}}{\cancel{\text{mol}}}\right) + \left(3 \cancel{\text{mol}} \times 32.0 \frac{\text{g}}{\cancel{\text{mol}}}\right) \longrightarrow$$

$$\left(2 \cancel{\text{mol}} \times 64.1 \frac{\text{g}}{\cancel{\text{mol}}}\right) + \left(2 \cancel{\text{mol}} \times 18.0 \frac{\text{g}}{\cancel{\text{mol}}}\right)$$

$68.2 \text{ g } H_2S + 96.0 \text{ g } O_2 \longrightarrow 128.2 \text{ g } SO_2 + 36.0 \text{ g } H_2O$

$164.2 \text{ g} = 164.2 \text{ g}$

3. Interpret the equation for the formation of water from its elements in terms of numbers of molecules and moles, and volumes of gases at STP.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

4. Balance the following equation:

$$C_2H_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Interpret the balanced equation in terms of relative numbers of moles, volumes of gas at STP, and masses of reactants and products.

Figure 12.2 summarizes the information derived from the balanced chemical equation for the formation of ammonia. As you can see, the mass of the reactants equals the mass of the products. In addition, the number of atoms of each type in the reactants equals the number of atoms of each type in the product. Mass and atoms are conserved in every chemical reaction. However, molecules, formula units, moles, and volumes are not necessarily conserved—although they may be. Consider, for example, the formation of hydrogen iodide.

$$H_2(g) + I_2(g) \longrightarrow 2HI(g)$$

In this reaction, molecules, moles, and volume are all conserved. But in the majority of chemical reactions, they are not.

$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
2 atoms N	+	6 atoms H	$\longrightarrow$	2 atoms N and 6 atoms H
1 molecule N_2	+	3 molecules H_2	$\longrightarrow$	2 molecules NH_3
10 molecules N_2	+	30 molecules H_2	$\longrightarrow$	20 molecules NH_3
$1 \times$ 6.02×10^{23} molecules N_2	+	$3 \times$ 6.02×10^{23} molecules H_2	$\longrightarrow$	$2 \times$ 6.02×10^{23} molecules NH_3
1 mol N_2	+	3 mol H_2	$\longrightarrow$	2 mol NH_3
28.0 g N_2	+	3×2.0 g H_2	$\longrightarrow$	2×17.0 g NH_3
		34.0 g reactants	$\longrightarrow$	34.0 g products
Assume STP 22.4 L	+	22.4 L 22.4 L 22.4 L	$\longrightarrow$	22.4 L 22.4 L
22.4 L N_2		67.2 L H_2		44.8 L NH_3

Figure 12.2 Interpreting a Balanced Chemical Equation
The balanced chemical equation for the formation of ammonia can be interpreted in several ways.
Predict *How many molecules of NH_3 could be made from 5 molecules of N_2 and 15 molecules of H_2?*

See balancing chemical equations animated online. KINETIC ART

12.1 LessonCheck

ONLINE PROBLEMS

5. Explain How do chemists use balanced equations?

6. Identify Chemical reactions can be described in terms of what quantities?

7. Explain How is a balanced equation similar to a recipe?

8. Identify What quantities are always conserved in chemical reactions?

9. Apply Concepts Interpret the given equation in terms of relative numbers of representative particles, numbers of moles, and masses of reactants and products.

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

10. Apply Concepts Balance this equation:

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

Show that the balanced equation obeys the law of conservation of mass.

12.2 Chemical Calculations

CHEMISTRY & YOU

Q: *How do manufacturers know how to make enough of their desired product?* Chemical plants produce ammonia by combining nitrogen with hydrogen. If too much ammonia is produced, then it might be wasted. But if too little is produced, then there might not be enough for all their customers. In this lesson, you will learn how to use a balanced chemical equation to calculate the amount of product formed in a chemical reaction.

Key Questions

How are mole ratios used in chemical calculations?

What is the general procedure for solving a stoichiometric problem?

Vocabulary

- mole ratio

Writing and Using Mole Ratios

How are mole ratios used in chemical calculations?

As you learned in the previous lesson, a balanced chemical equation provides a great deal of quantitative information. It relates particles (atoms, molecules, formula units), moles of substances, and masses. A balanced chemical equation also is essential for all calculations involving amounts of reactants and products. For example, suppose you know the number of moles of one substance. The balanced chemical equation allows you to determine the number of moles of all other substances in the reaction.

Look again at the balanced equation for the production of ammonia.

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

The most important interpretation of this equation is that 1 mol of nitrogen reacts with 3 mol of hydrogen to form 2 mol of ammonia. Based on this interpretation, you can write ratios that relate moles of reactants to moles of product. A **mole ratio** is a conversion factor derived from the coefficients of a balanced chemical equation interpreted in terms of moles. **In chemical calculations, mole ratios are used to convert between a given number of moles of a reactant or product to moles of a different reactant or product.** Three mole ratios derived from the balanced equation above are

$$\frac{1\ \text{mol } N_2}{3\ \text{mol } H_2} \quad \frac{2\ \text{mol } NH_3}{1\ \text{mol } N_2} \quad \frac{3\ \text{mol } H_2}{2\ \text{mol } NH_3}$$

Mole-Mole Calculations In the mole ratio below, *W* is the unknown, wanted, quantity and *G* is the given quantity. The values of *a* and *b* are the coefficients from the balanced equation. Thus, a general solution for a mole-mole problem, such as Sample Problem 12.3, is given by

$$\underbrace{x\ \cancel{\text{mol } G}}_{\text{Given}} \times \underbrace{\frac{b\ \text{mol } W}{a\ \cancel{\text{mol } G}}}_{\text{Mole ratio}} = \underbrace{\frac{xb}{a}\ \text{mol } W}_{\text{Calculated}}$$

Sample Problem 12.3

Calculating Moles of a Product

How many moles of NH_3 are produced when 0.60 mol of nitrogen reacts with hydrogen?

1 Analyze **List the known and the unknown.** The conversion is mol $N_2 \longrightarrow$ mol NH_3. According to the balanced equation, 1 mol N_2 combines with 3 mol H_2 to produce 2 mol NH_3. To determine the number of moles of NH_3, the given quantity of N_2 is multiplied by the form of the mole ratio from the balanced equation that allows the given unit to cancel.

KNOWN

moles of nitrogen = 0.60 mol N_2

UNKNOWN

moles of ammonia = ? mol NH_3

2 Calculate **Solve for the unknown.**

Write the mole ratio that will allow you to convert from moles N_2 to moles NH_3.

$$\frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2}$$

Multiply the given quantity of N_2 by the mole ratio in order to find the moles of NH_3.

$$0.60 \cancel{\text{mol } N_2} \times \frac{2 \text{ mol } NH_3}{1 \cancel{\text{mol } N_2}} = 1.2 \text{ mol } NH_3$$

3 Evaluate **Does the result make sense?** The ratio of 1.2 mol NH_3 to 0.60 mol N_2 is 2:1, as predicted by the balanced equation.

Remember that the mole ratio must have N_2 on the bottom so that the mol N_2 in the mol ratio will cancel with mol N_2 in the known.

11. This equation shows the formation of aluminum oxide, which is found on the surface of aluminum objects exposed to the air.

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

a. Write the six mole ratios that can be derived from this equation.

b. How many moles of aluminum are needed to form 3.7 mol Al_2O_3?

12. According to the equation in Problem 11,

a. How many moles of oxygen are required to react completely with 14.8 mol Al?

b. How many moles of Al_2O_3 are formed when 0.78 mol O_2 reacts with aluminum?

Mass-Mass Calculations No laboratory balance can measure substances directly in moles. Instead, the amount of a substance is usually determined by measuring its mass in grams. From the mass of a reactant or product, the mass of any other reactant or product in a given chemical equation can be calculated. The mole interpretation of a balanced equation is the basis for this conversion. If the given sample is measured in grams, then the mass can be converted to moles by using the molar mass. Then the mole ratio from the balanced equation can be used to calculate the number of moles of the unknown. If it is the mass of the unknown that needs to be determined, the number of moles of the unknown can be multiplied by the molar mass. As in mole-mole calculations, the unknown can be either a reactant or a product.

Figure 12.3 Ammonia in Space
In this Hubble Space Telescope image, clouds of condensed ammonia are visible covering the surface of Saturn.

Steps for Solving a Mass-Mass Problem Mass-mass problems are solved in basically the same way as mole-mole problems. The steps for the mass-mass conversion of any given mass (*G*) to any wanted mass (*W*) are outlined below.

1. Change the mass of G to moles of G (mass $G \longrightarrow$ mol G) by using the molar mass of G.

$$\text{mass } G \times \frac{1 \text{ mol } G}{\text{molar mass } G} = \text{mol } G$$

2. Change the moles of G to moles of *W* (mol $G \longrightarrow$ mol *W*) by using the mole ratio from the balanced equation.

$$\text{mol } G \times \frac{b \text{ mol } W}{a \text{ mol } G} = \text{mol } W$$

3. Change the moles of *W* to grams of *W* (mol $W \longrightarrow$ mass *W*) by using the molar mass of *W*.

$$\text{mol } W \times \frac{\text{molar mass } W}{1 \text{ mol } W} = \text{mass } W$$

Figure 12.4 shows another way to represent the steps for doing mole-mass and mass-mole stoichiometric calculations. For a mole-mass problem, the first conversion (from mass to moles) is skipped. For a mass-mole problem, the last conversion (from moles to mass) is skipped. You can use parts of the three-step process shown in Figure 12.4 as they are appropriate to the problem you are solving.

Figure 12.4 Mass-Mass Conversion Steps
This general solution diagram indicates the steps necessary to solve a mass-mass stoichiometry problem: Convert mass to moles, use the mole ratio, and then convert moles to mass.
Infer ***Is the given always a reactant?***

$$aG \longrightarrow bW$$

(given quantity) (wanted quantity)

$$\text{mass of } G \times \boxed{\frac{1 \text{ mol } G}{\text{mass } G}} \rightarrow \text{mol } G \times \boxed{\frac{b \text{ mol } W}{a \text{ mol } G}} \rightarrow \text{mol } W \times \boxed{\frac{\text{mass } W}{1 \text{ mol } W}} \rightarrow \text{mass of } W$$

Mass-mole conversion — Mole ratio from balanced equation — Mole-mass conversion

Sample Problem 12.4

Calculating the Mass of a Product

Ammonia (NH_3) clouds are present around some planets, as in Figure 12.3. Calculate the number of grams of NH_3 produced by the reaction of 5.40 g of hydrogen with an excess of nitrogen. The balanced equation is

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

1 Analyze List the knowns and the unknown. The mass of hydrogen will be used to find the mass of ammonia: g $H_2 \longrightarrow$ g NH_3. The coefficients of the balanced equation show that 3 mol H_2 reacts with 1 mol N_2 to produce 2 mol NH_3. The following steps are necessary to determine the mass of ammonia:

$$\text{g } H_2 \longrightarrow \text{mol } H_2 \longrightarrow \text{mol } NH_3 \longrightarrow \text{g } NH_3$$

KNOWNS

mass of hydrogen = 5.40 g H_2

2 mol NH_3/3 mol H_2 (from balanced equation)

1 mol H_2 = 2.0 g H_2 (molar mass)

1 mol NH_3 = 17.0 g NH_3 (molar mass)

UNKNOWN

mass of ammonia = ? g NH_3

2 Calculate Solve for the unknown.

Start with the given quantity, and convert from mass to moles.

$$5.40\,\cancel{\text{g H}_2} \times \frac{1 \text{ mol H}_2}{2.0\,\cancel{\text{g H}_2}}$$

Don't forget to cancel the units at each step.

Then convert from moles of reactant to moles of product by using the correct mole ratio.

$$5.40\,\cancel{\text{g H}_2} \times \frac{1\,\cancel{\text{mol H}_2}}{2.0\,\cancel{\text{g H}_2}} \times \frac{2 \text{ mol NH}_3}{3\,\cancel{\text{mol H}_2}}$$

$$\text{g } H_2 \longrightarrow \text{mol } H_2 \longrightarrow \text{mol } NH_3 \longrightarrow \text{g } NH_3$$

Finish by converting from moles to grams. Use the molar mass of NH_3.

$$\underbrace{5.40\,\cancel{\text{g H}_2}}_{\text{Given quantity}} \times \underbrace{\frac{1\,\cancel{\text{mol H}_2}}{2.0\,\cancel{\text{g H}_2}}}_{\text{Change given unit to moles}} \times \underbrace{\frac{2\,\cancel{\text{mol NH}_3}}{3\,\cancel{\text{mol H}_2}}}_{\text{Mole ratio}} \times \underbrace{\frac{17.0 \text{ g NH}_3}{1\,\cancel{\text{mol NH}_3}}}_{\text{Change moles to grams}} = 31 \text{ g NH}_3$$

3 Evaluate Does the result make sense? Because there are three conversion factors involved in this solution, it is more difficult to estimate an answer. However, because the molar mass of NH_3 is substantially greater than the molar mass of H_2, the answer should have a larger mass than the given mass. The answer should have two significant figures.

13. Acetylene gas (C_2H_2) is produced by adding water to calcium carbide (CaC_2).

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

How many grams of acetylene are produced by adding water to 5.00 g CaC_2?

14. Use the equation in Question 13 to determine how many moles of CaC_2 are needed to react completely with 49.0 g H_2O.

Other Stoichiometric Calculations

What is the general procedure for solving a stoichiometric problem?

As you already know, you can obtain mole ratios from a balanced chemical equation. From the mole ratios, you can calculate any measurement unit that is related to the mole. The given quantity can be expressed in numbers of representative particles, units of mass, or volumes of gases at STP. The problems can include mass-volume, particle-mass, and volume-volume calculations. For example, you can use stoichiometry to relate volumes of reactants and products in the reaction shown in Figure 12.5. **In a typical stoichiometric problem, the given quantity is first converted to moles. Then, the mole ratio from the balanced equation is used to calculate the number of moles of the wanted substance. Finally, the moles are converted to any other unit of measurement related to the unit mole, as the problem requires.**

CHEMISTRY & YOU

Q: *How do you think air bag manufacturers know how to get the right amount of air in an inflated air bag?*

Thus far, you have learned how to use the relationship between moles and mass (1 mol = molar mass) in solving mass-mass, mass-mole, and mole-mass stoichiometric problems. The mole-mass relationship gives you two conversion factors.

$$\frac{1\ \text{mol}}{\text{molar mass}} \quad \text{and} \quad \frac{\text{molar mass}}{1\ \text{mol}}$$

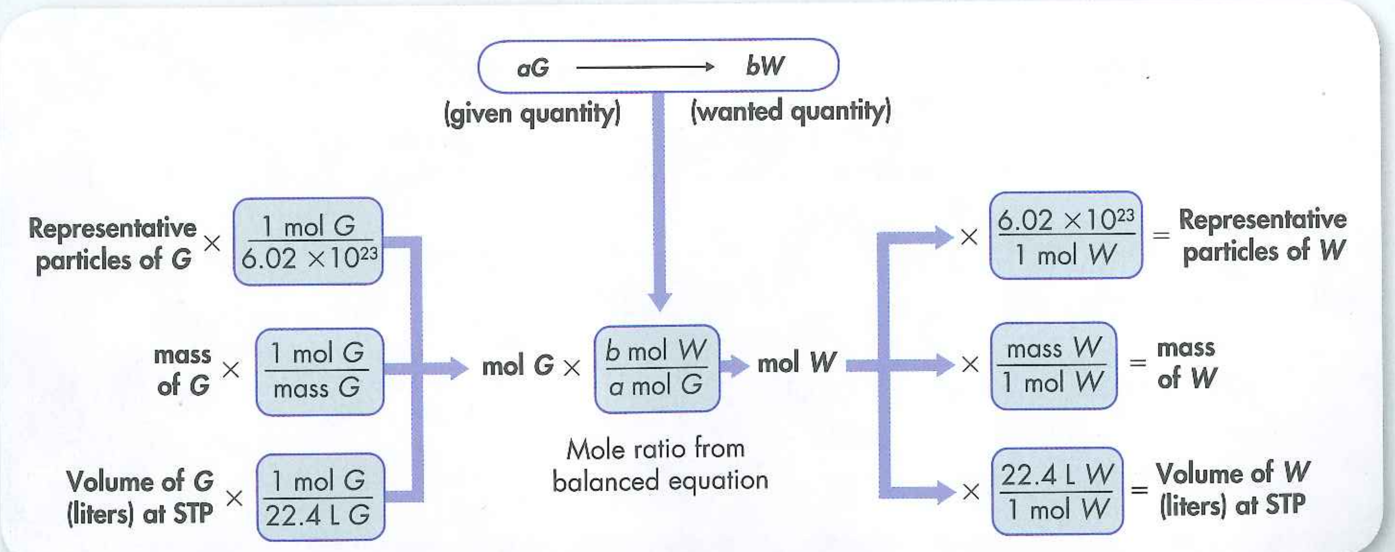
Recall from Chapter 10 that the mole can be related to other quantities as well. For example, 1 mol = 6.02×10^{23} representative particles, and 1 mol of a gas = 22.4 L at STP. These two relationships provide four more conversion factors that you can use in stoichiometric calculations.

$$\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{particles}} \quad \text{and} \quad \frac{6.02 \times 10^{23}\ \text{particles}}{1\ \text{mol}}$$

$$\frac{1\ \text{mol}}{22.4\ \text{L}} \quad \text{and} \quad \frac{22.4\ \text{L}}{1\ \text{mol}}$$

Figure 12.5 summarizes the steps for a typical stoichiometric problem. Notice that the units of the given quantity will not necessarily be the same as the units of the wanted quantity. For example, given the mass of G, you might be asked to calculate the volume of W at STP.

Figure 12.5
Solving Stoichiometric Problems
With your knowledge of conversion factors and this problem-solving approach, you can solve a variety of stoichiometric problems.
Identify *What conversion factor is used to convert moles to representative particles?*



Sample Problem 12.5

Calculating Molecules of a Product

How many molecules of oxygen are produced when 29.2 g of water is decomposed by electrolysis according to this balanced equation?

$$2H_2O(l) \xrightarrow{\text{electricity}} 2H_2(g) + O_2(g)$$

1 Analyze **List the knowns and the unknown.** The following calculations need to be performed:

$$\text{g } H_2O \longrightarrow \text{mol } H_2O \longrightarrow \text{mol } O_2 \longrightarrow \text{molecules } O_2$$

The appropriate mole ratio relating mol O_2 to mol H_2O from the balanced equation is 1 mole O_2/2 mol H_2O.

KNOWNS

mass of water = 29.2 g H_2O

1 mol O_2/2 mol H_2O (from balanced equation)

1 mol H_2O = 18.0 g H_2O (molar mass)

1 mol O_2 = 6.02×10^{23} molecules O_2

UNKNOWN

molecules of oxygen = ? moleciules O_2

2 Calculate **Solve for the unknown.**

Start with the given quantity, and convert from mass to moles.

$$29.2\ \cancel{\text{g } H_2O} \times \frac{1 \text{ mol } H_2O}{18.0\ \cancel{\text{g } H_2O}}$$

Remember to start your calculations with the given quantity, even if the given quantity is a product in the reaction.

Then, convert from moles of reactant to moles of product.

$$29.2\ \cancel{\text{g } H_2O} \times \frac{1\ \cancel{\text{mol } H_2O}}{18.0\ \cancel{\text{g } H_2O}} \times \frac{1 \text{ mol } O_2}{2\ \cancel{\text{mol } H_2O}}$$

Finish by converting from moles to molecules.

$$\underbrace{29.2\ \cancel{\text{g } H_2O}}_{\text{Given quantity}} \times \underbrace{\frac{1\ \cancel{\text{mol } H_2O}}{18.0\ \cancel{\text{g } H_2O}}}_{\text{Change to moles}} \times \underbrace{\frac{1\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } H_2O}}}_{\text{Mole ratio}} \times \underbrace{\frac{6.02 \times 10^{23} \text{ molecules } O_2}{1\ \cancel{\text{mol } O_2}}}_{\text{Change to molecules}}$$

$$= 4.88 \times 10^{23} \text{ molecules } O_2$$

3 Evaluate **Does the result make sense?** The given mass of water should produce a little less than 1 mol of oxygen, or a little less than Avogadro's number of molecules. The answer should have three significant figures.

15. How many molecules of oxygen are produced by the decomposition of 6.54 g of potassium chlorate ($KClO_3$)?

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

16. The last step in the production of nitric acid is the reaction of nitrogen dioxide with water.

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$$

How many grams of nitrogen dioxide must react with water to produce 5.00×10^{22} molecules of nitrogen monoxide?

Sample Problem 12.6

Volume-Volume Stoichiometric Calculations

Nitrogen monoxide and oxygen gas combine to form the brown gas nitrogen dioxide, which contributes to photochemical smog. How many liters of nitrogen dioxide are produced when 34 L of oxygen react with an excess of nitrogen monoxide? Assume conditions are at STP.

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

KNOWNS

volume of oxygen = 34 L O_2
2 mol NO_2/1 mol O_2 (from balanced equation)
1 mol O_2 = 22.4 L O_2 (at STP)
1 mol NO_2 = 22.4 L NO_2 (at STP)

UNKNOWN

volume of nitrogen dioxide = ? L NO_2

1 Analyze **List the knowns and the unknown.** The following calculations need to be performed:

$$\text{L } O_2 \longrightarrow \text{mol } O_2 \longrightarrow \text{mol } NO_2 \longrightarrow \text{L } NO_2$$

For gaseous reactants and products at STP, 1 mol of a gas has a volume of 22.4 L.

2 Calculate **Solve for the unknown.**

Start with the given quantity, and convert from volume to moles by using the mole-volume ratio.

$$34\,\cancel{\text{L } O_2} \times \frac{1 \text{ mol } O_2}{22.4\,\cancel{\text{L } O_2}}$$

Then, convert from moles of reactant to moles of product by using the correct mole ratio.

$$34\,\cancel{\text{L } O_2} \times \frac{1\,\cancel{\text{mol } O_2}}{22.4\,\cancel{\text{L } O_2}} \times \frac{2 \text{ mol } NO_2}{1\,\cancel{\text{mol } O_2}}$$

Finish by converting from moles to liters. Use the mole-volume ratio.

$$\underbrace{34\,\cancel{\text{L } O_2}}_{\text{Given quantity}} \times \underbrace{\frac{1\,\cancel{\text{mol } O_2}}{22.4\,\cancel{\text{L } O_2}}}_{\text{Change to moles}} \times \underbrace{\frac{2\,\cancel{\text{mol } NO_2}}{1\,\cancel{\text{mol } O_2}}}_{\text{Mole ratio}} \times \underbrace{\frac{22.4 \text{ L } NO_2}{1\,\cancel{\text{mol } NO_2}}}_{\text{Change to liters}} = \boxed{68 \text{ L } NO_2}$$

3 Evaluate **Does the result make sense?** Because 2 mol NO_2 are produced for each 1 mol O_2 that reacts, the volume of NO_2 should be twice the given volume of O_2. The answer should have two significant figures.

17. The equation for the combustion of carbon monoxide is

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

How many liters of oxygen are required to burn 3.86 L of carbon monoxide?

18. Phosphorus and hydrogen can be combined to form phosphine (PH_3).

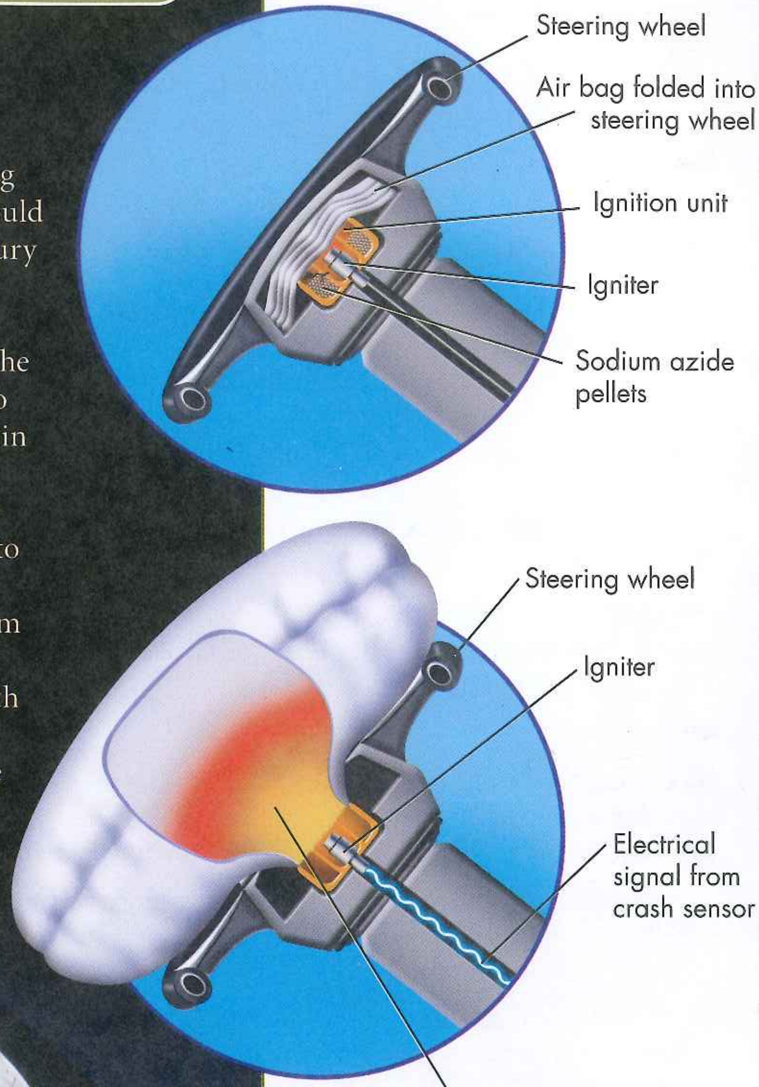
$$P_4(s) + 6H_2(g) \longrightarrow 4PH_3(g)$$

How many liters of phosphine are formed when 0.42 L of hydrogen reacts with phosphorus?

Stoichiometric Safety

In a car collision, proper inflation of an air bag may save your life. Too much air in the bag could make the bag too hard, which could cause injury because the bag wouldn't effectively cushion the blow. Too little air in the bag could be insufficient to prevent a driver's impact with the steering wheel. Engineers use stoichiometry to determine the exact quantity of each reactant in the air bag's inflation system.

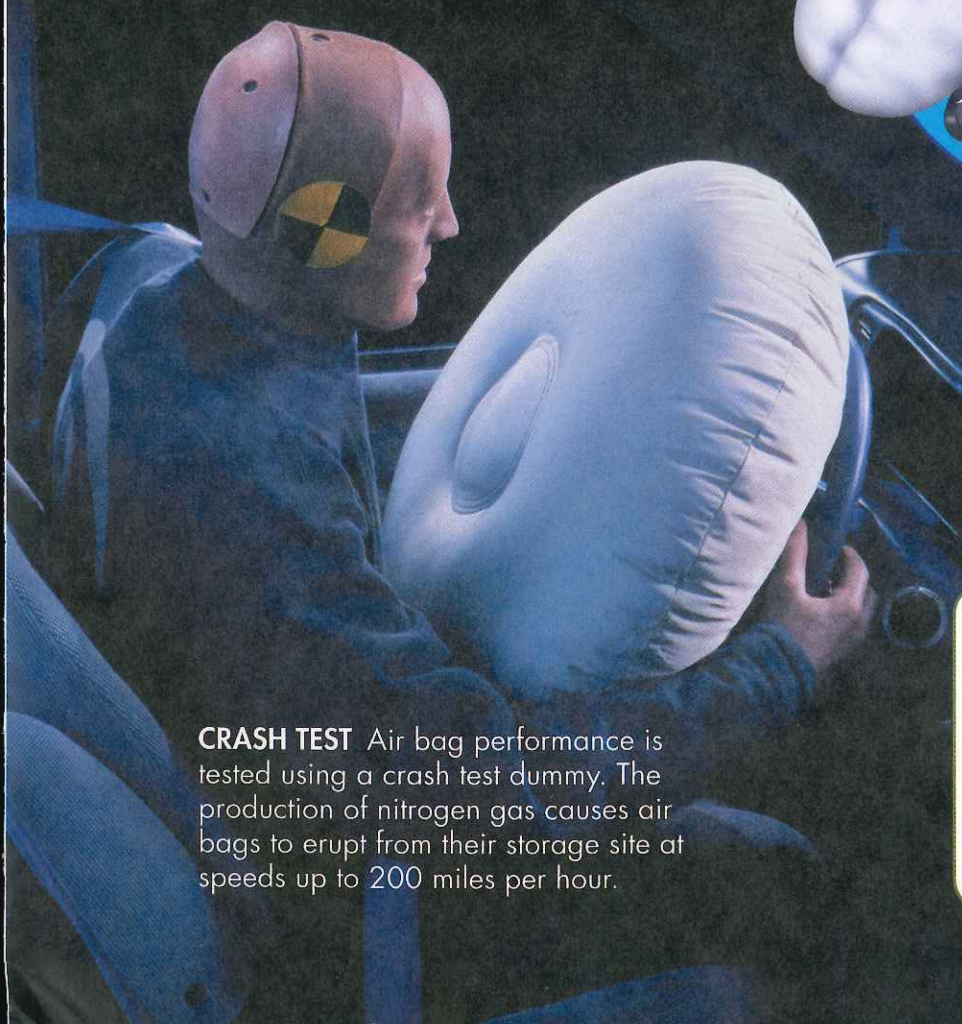
When a crash occurs, a series of reactions happen. Sodium azide (NaN_3) decomposes into sodium metal and nitrogen gas. The nitrogen gas causes the air bag to inflate, but the sodium can react explosively with water. So, air bags contain potassium nitrate (KNO_3) to react with the sodium. Silicon dioxide is also included in the air bag to react with the products of the second reaction. This final reaction produces a harmless substance.



Sodium azide pellets decomposing

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

$10Na(s) + 2KNO_3(s) \longrightarrow K_2O(s) + 5Na_2O(s) + N_2(g)$



CRASH TEST Air bag performance is tested using a crash test dummy. The production of nitrogen gas causes air bags to erupt from their storage site at speeds up to 200 miles per hour.

Take It Further

1. Draw Conclusions If a reaction in an air bag does not occur as intended, how might this affect the performance of an air bag?

2. Explain Research the regulations on automotive air bags, and explain why air bags are not safe for all passengers.

SampleProblem 12.7

Finding the Volume of a Gas Needed for a Reaction

Assuming STP, how many milliliters of oxygen are needed to produce 20.4 mL SO_3 according to this balanced equation?

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

1 Analyze **List the knowns and the unknown.** For a reaction involving gaseous reactants or products, the coefficients also indicate relative amounts of each gas. So, you can use volume ratios in the same way you have used mole ratios.

KNOWNS

volume of sulfur trioxide = 20.4 mL

1 ml O_2/2 ml SO_3 (from balanced equation)

UNKNOWN

volume of oxygen = ? mL O_2

2 Calculate **Solve for the unknown.**

Multiply the given volume by the appropriate volume ratio.

$$20.4\ \cancel{\text{mL}\ SO_3} \times \frac{1\ \text{mL}\ O_2}{2\ \cancel{\text{mL}\ SO_3}} = \boxed{10.2\ \text{mL}\ O_2}$$

The volume ratio can be written using milliliters as the units instead of liters.

3 Evaluate **Does the result make sense?** Because the volume ratio is 2 volumes SO_3 to 1 volume O_2, the volume of O_2 should be half the volume of SO_3. The answer should have three significant figures.

Use the following chemical equation to answer Problems 19 and 20.

$$CS_2(l) + 3O_2(g) \longrightarrow CO_2(g) + 2SO_2(g)$$

19. Calculate the volume of sulfur dioxide, in milliliters, produced when 27.9 mL O_2 reacts with carbon disulfide.

20. How many deciliters of carbon dioxide are produced when 0.38 L SO_2 is formed?

ONLINE PROBLEMS 12.2 LessonCheck

21. **Explain** How are mole ratios used in chemical calculations?

22. **Sequence** Outline the sequence of steps needed to solve a typical stoichiometric problem.

23. **Calculate** The combustion of acetylene gas is represented by this equation:

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

a. How many grams of CO_2 and grams of H_2O are produced when 52.0 g C_2H_2 burn in oxygen?

b. How many moles of H_2O are produced when 64.0 g C_2H_2 burn in oxygen?

24. **Apply Concepts** Write the 12 mole ratios that can be derived from the equation for the combustion of isopropyl alcohol.

$$2C_3H_7OH(l) + 9O_2(g) \longrightarrow 6CO_2(g) + 8H_2O(g)$$

BIGIDEA
THE MOLE AND QUANTIFYING MATTER

25. Use what you have learned about stoichiometric calculations to explain the following statement: Stoichiometric calculations are not possible without a balanced chemical equation.

Small-Scale Lab

Analysis of Baking Soda

(Probe or sensor version of this lab is available in the *Probeware Lab Manual.*)

Purpose

To determine the mass of sodium hydrogen carbonate in a sample of baking soda, using stoichiometry

Materials

- baking soda
- 3 plastic cups
- soda straw
- balance
- pipets of HCl, NaOH, and thymol blue

Procedure

A. Measure the mass of a clean, dry plastic cup.

B. Using the straw as a scoop, fill one end with baking soda to a depth of about 1 cm. Add the sample to the cup and measure its mass again.

C. Place two HCl pipets that are about 3/4 full into a clean cup and measure the mass of the system.

D. Transfer the contents of both HCl pipets to the cup containing baking soda. Swirl until the fizzing stops. Wait 5–10 minutes to be sure the reaction is complete. Measure the mass of the two empty HCl pipets in their cup again.

E. Add 5 drops of thymol blue to the plastic cup.

F. Place two full NaOH pipets in a clean cup and measure the mass of the system.

G. Add NaOH slowly to the baking soda/HCl mixture until the pink color just disappears. Measure the mass of the NaOH pipets in their cup again.

Analyze

Using your experimental data, record the answers to the following questions below your data table.

1. Evaluate Write a balanced equation for the reaction between baking soda ($NaHCO_3$) and HCl.

2. Calculate Calculate the mass in grams of the baking soda.

(Step B − Step A)

3. Calculate Calculate the total mmol of 1*M* HCl. *Note:* Every gram of HCl contains 1 mmol.

(Step C − Step D) × 1.00 mmol/g

4. Calculate Calculate the total mmol of 0.5*M* NaOH. *Note:* Every gram of NaOH contains 0.5 mmol.

(Step F − Step G) × 0.500 mmol/g

5. Calculate Calculate the mmol of HCl that reacted with the baking soda. *Note:* The NaOH measures the amount of HCl that did not react.

(Step 3 − Step 4)

6. Calculate Calculate the mass of the baking soda from the reaction data.

(0.084 g/mmol × Step 5)

7. Calculate Calculate the percent error of the experiment.

$$\frac{(\text{Step 2} - \text{Step 6})}{\text{Step 2}} \times 100\%$$

You're the Chemist

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. Analyze Data For each calculation you did, substitute each quantity (number and unit) into the equation and cancel the units to explain why each step gives the quantity desired.

2. Design an Experiment Baking powder consists of a mixture of baking soda, sodium hydrogen carbonate, and a solid acid, usually calcium dihydrogen phosphate ($Ca(H_2PO_4)_2$). Design and carry out an experiment to determine the percentage of baking soda in baking powder.

12.3 Limiting Reagent and Percent Yield

CHEMISTRY & YOU

Q: *What determines how much product you can make?* If a carpenter had two tabletops and seven table legs, he would have difficulty building more than one functional four-legged table. The first table would require four of the legs, leaving just three legs for the second table. In this case, the number of table legs is limiting the construction of four-legged tables. In this lesson you will learn how the amount of product is limited in a chemical reaction.

Key Questions

How is the amount of product in a reaction affected by an insufficient quantity of any of the reactants?

What does the percent yield of a reaction measure?

Vocabulary

- limiting reagent
- excess reagent
- theoretical yield
- actual yield
- percent yield

Limiting and Excess Reagents

How is the amount of product in a reaction affected by an insufficient quantity of any of the reactants?

Many cooks follow a recipe when making a new dish. They know that sufficient quantities of all the ingredients must be available in order to follow the recipe. Suppose, for example, that you are preparing to make tacos like the ones in Figure 12.6. You would have more than enough meat, cheese, lettuce, tomatoes, sour cream, salsa, and seasoning on hand. However, you have only two taco shells. The quantity of taco shells you have will limit the number of tacos you can make. Thus, the taco shells are the limiting ingredient in this cooking venture. A chemist often faces a similar situation. **In a chemical reaction, an insufficient quantity of any of the reactants will limit the amount of product that forms.**

Figure 12.6 Limiting Ingredients
The amount of product is determined by the quantity of the limiting reagent. In this example, the taco shells are the limiting reagent. No matter how much of the other ingredients you have, with two taco shells you can make only two tacos.

Chemical Equations

	$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
Microscopic recipe:	1 molecule N_2	+	3 molecules H_2	$\longrightarrow$	2 molecules NH_3
Macroscopic recipe:	1 mol N_2	+	3 mol H_2	$\longrightarrow$	2 mol NH_3

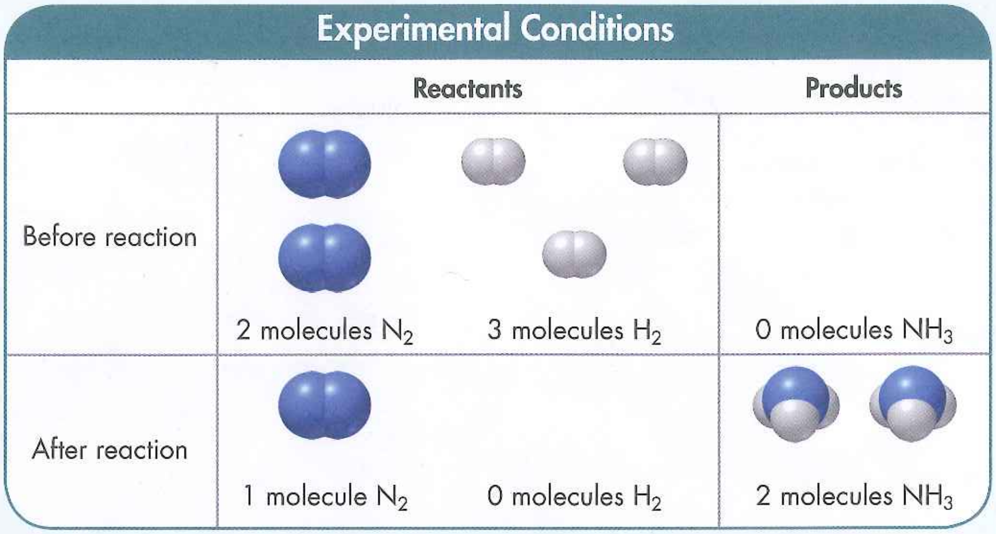


Figure 12.7 Limiting Reagent The "recipe" calls for three molecules of H_2 for every one molecule of N_2.

In this particular experiment, H_2 is the limiting reagent and N_2 is in excess.

Infer ***How would the amount of products formed change if you started with four molecules of N_2 and three molecules of H_2?***

As you know, a balanced chemical equation is a chemist's recipe. You can interpret the recipe on a microscopic scale (interacting particles) or on a macroscopic scale (interacting moles). The coefficients used to write the balanced equation give both the ratio of representative particles and the mole ratio. Recall the equation for the preparation of ammonia:

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

When one molecule (mole) of N_2 reacts with three molecules (moles) of H_2, two molecules (moles) of NH_3 are produced. What would happen if two molecules (moles) of N_2 reacted with three molecules (moles) of H_2? Would more than two molecules (moles) of NH_3 be formed? Figure 12.7 shows both the particle and the mole interpretations of this problem.

Before the reaction takes place, nitrogen and hydrogen are present in a 2:3 molecule (mole) ratio. The reaction takes place according to the balanced equation. One molecule (mole) of N_2 reacts with three molecules (moles) of H_2 to produce two molecules (moles) of NH_3. At this point, all the hydrogen has been used up, and the reaction stops. One molecule (mole) of unreacted nitrogen is left in addition to the two molecules (moles) of NH_3 that have been produced by the reaction.

In this reaction, only the hydrogen is completely used up. This reactant is the **limiting reagent**, or the reactant that determines the amount of product that can be formed by a reaction. The reaction occurs only until the limiting reagent is used up. By contrast, the reactant that is not completely used up in a reaction is called the **excess reagent**. In this example, nitrogen is the excess reagent because some nitrogen remains unreacted.

Sometimes in stoichiometric problems, the given quantities of reactants are expressed in units other than moles. In such cases, the first step in the solution is to convert the quantity of each reactant to moles. Then the limiting reagent can be identified. The amount of product formed in a reaction can be determined from the given amount of limiting reagent.

See limiting reagents animated online. KINETIC ART

CHEMISTRY & YOU

Q: *What determines how much product you can make in a chemical reaction?*

CHEM TUTOR

Sample Problem 12.8

Determining the Limiting Reagent in a Reaction

Copper reacts with sulfur to form copper(I) sulfide according to the following balanced equation:

$$2Cu(s) + S(s) \longrightarrow Cu_2S(s)$$

What is the limiting reagent when 80.0 g Cu reacts with 25.0 g S?

1 Analyze **List the knowns and the unknown.** The number of moles of each reactant must first be found. The balanced equation is used to calculate the number of moles of one reactant needed to react with the given amount of the other reactant.

KNOWNS

mass of copper = 80.0 g Cu

mass of sulfur = 25.0 g S

1 mol S / 2 mol Cu

UNKNOWN

limiting reagent = ?

2 Calculate **Solve for the unknown.**

Start with one of the reactants and convert from mass to moles.

$$80.0\,\cancel{g\,Cu} \times \frac{1\text{ mol Cu}}{63.5\,\cancel{g\,Cu}} = 1.26\text{ mol Cu}$$

Then, convert the mass of the other reactant to moles.

$$25.0\,\cancel{g\,S} \times \frac{1\text{ mol S}}{32.1\,\cancel{g\,S}} = 0.779\text{ mol S}$$

Now convert moles of Cu to moles of S needed to react with 1.26 moles of Cu.

$$\underbrace{1.26\,\cancel{mol\,Cu}}_{\text{Given quantity}} \times \underbrace{\frac{1\text{ mol S}}{2\,\cancel{mol\,Cu}}}_{\text{Mole ratio}} = \underbrace{0.630\text{ mol S}}_{\text{Needed amount}}$$

It doesn't matter which reactant you use. If you used the actual number of moles of S to find the amount of copper needed, then you would still identify copper as the limiting reagent.

Compare the amount of sulfur needed with the given amount of sulfur.

0.630 mol S (amount needed to react) $<$ 0.779 mol S (given amount)
Sulfur is in excess, so copper is the limiting reagent.

3 Evaluate **Do the results make sense?** Since the ratio of the given mol Cu to mol S was less than the ratio (2:1) from the balanced equation, copper should be the limiting reagent.

26. The equation for the complete combustion of ethene (C_2H_4) is

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

If 2.70 mol C_2H_4 reacts with 6.30 mol O_2, identify the limiting reagent.

27. Hydrogen gas can be produced by the reaction of magnesium metal with hydrochloric acid.

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Identify the limiting reagent when 6.00 g HCl reacts with 5.00 g Mg.

In Sample Problem 12.8, you may have noticed that even though the mass of copper used in the reaction is greater than the mass of sulfur, copper is the limiting reagent. The reactant that is present in the smaller amount by mass or volume is not necessarily the limiting reagent.

SampleProblem 12.9

Using a Limiting Reagent to Find the Quantity of a Product

What is the maximum number of grams of Cu_2S that can be formed when 80.0 g Cu reacts with 25.0 g S?

$$2Cu(s) + S(s) \longrightarrow Cu_2S(s)$$

1 Analyze List the knowns and the unknown. The limiting reagent, which was determined in Sample Problem 12.8, is used to calculate the maximum amount of Cu_2S formed.

$$\text{mol Cu} \longrightarrow \text{mol } Cu_2S \longrightarrow \text{g } Cu_2S$$

KNOWNS

limiting reagent = 1.26 mol Cu (from sample problem 12.8)
1 mol Cu_2S = 159.1 g Cu_2S (molar mass)
1 mol Cu_2S / 2 mol Cu (from balanced equation)

UNKNOWN

yield = ? g Cu_2S

2 Calculate Solve for the unknown.

Start with the moles of the limiting reagent and convert to moles of the product. Use the mole ratio from the balanced equation.

$$1.26\ \cancel{\text{mol Cu}} \times \frac{1\ \text{mol } Cu_2S}{2\ \cancel{\text{mol Cu}}}$$

Finish the calculation by converting from moles to mass of product.

$$1.26\ \cancel{\text{mol Cu}} \times \frac{1\ \cancel{\text{mol } Cu_2S}}{2\ \cancel{\text{mol Cu}}} \times \frac{159.1\ \text{g } Cu_2S}{1\ \cancel{\text{mol } Cu_2S}} = \boxed{1.00 \times 10^2\ \text{g } Cu_2S}$$

3 Evaluate Do the results make sense? Copper is the limiting reagent in this reaction. The maximum number of grams of Cu_2S produced should be more than the amount of copper that initially reacted because copper is combining with sulfur. However, the mass of Cu_2S produced should be less than the total mass of the reactants (105.0 g) because sulfur was in excess.

28. The equation below shows the incomplete combustion of ethene.

$$C_2H_4(g) + 2O_2(g) \longrightarrow 2CO(g) + 2H_2O(g)$$

If 2.70 mol C_2H_4 is reacted with 6.30 mol O_2,

a. identify the limiting reagent.

b. calculate the moles of water produced.

29. The heat from an acetylene torch is produced by burning acetylene (C_2H_2) in oxygen.

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

How many grams of water can be produced by the reaction of 2.40 mol C_2H_2 with 7.40 mol O_2?

Quick Lab

Purpose To illustrate the concept of a limiting reagent in a chemical reaction

Materials

- graduated cylinder
- balance
- 3 250-mL Erlenmeyer flasks
- 3 rubber balloons
- 4.2 g magnesium ribbon
- 300 mL 1.0*M* hydrochloric acid

Limiting Reagents

Procedure

1. Add 100 mL of the hydrochloric acid solution to each flask.

2. Weigh out 0.6 g, 1.2 g, and 2.4 g of magnesium ribbon, and place each sample into its own balloon.

3. Stretch the end of each balloon over the mouth of each flask. Do not allow the magnesium ribbon in the balloon to fall into the flask.

4. Magnesium reacts with hydrochloric acid to form hydrogen gas. When you mix the magnesium with the hydrochloric acid in the next step, you will generate a certain volume of hydrogen gas. How do you think the volume of hydrogen produced in each flask will compare?

5. Lift up on each balloon and shake the magnesium into each flask. Observe the volume of gas produced until the reaction in each flask is completed. Record your observations.

Analyze and Conclude

1. Analyze Data How did the volumes of hydrogen gas produced, as measured by the size of the balloons, compare? Did the results agree with your prediction?

2. Apply Concepts Write a balanced equation for the reaction you observed.

3. Calculate The 100 mL of hydrochloric acid contained 0.10 mol HCl. Show by calculation why the balloon with 1.2 g Mg inflated to about twice the size of the balloon with 0.60 g Mg.

4. Calculate Show by calculation why the balloons with 1.2 g and 2.4 g Mg inflated to approximately the same volume. What was the limiting reagent when 2.4 g Mg was added to the acid?

Percent Yield

What does the percent yield of a reaction measure?

When a teacher gives an exam to the class, every student could get a grade of 100 percent. However, this outcome generally does not occur. Instead, the performance of the class is usually spread over a range of grades. Your exam grade, expressed as a percentage, is a ratio of two items. The first item is the number of questions you answered correctly. The second is the total number of questions. The grade compares how well you performed with how well you could have performed if you had answered all the questions correctly. Chemists perform similar calculations in the laboratory when the product from a chemical reaction is less than expected, based on the balanced chemical equation.

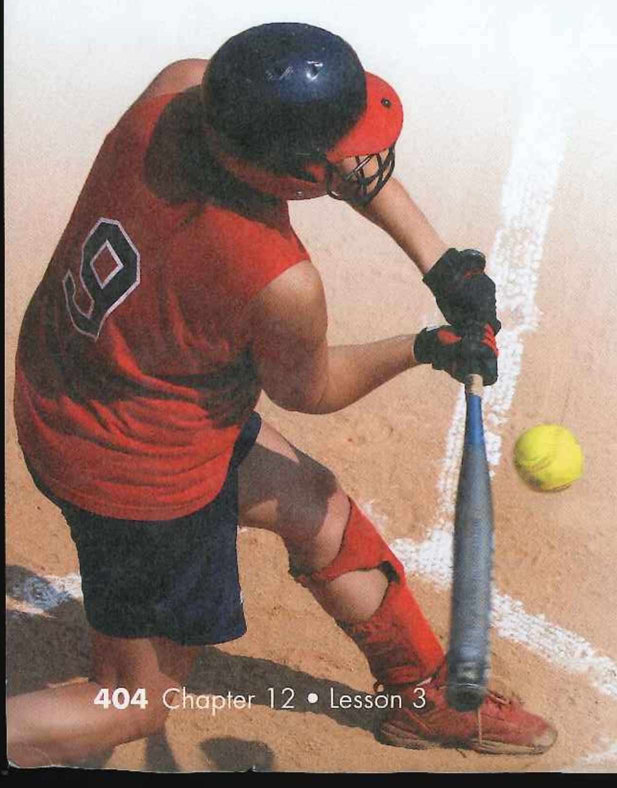


Figure 12.8 Batting Average
A batting average is actually a percent yield. A batting average is calculated by dividing the number of hits a batter has had (actual yield) by the number of at-bats (theoretical yield).

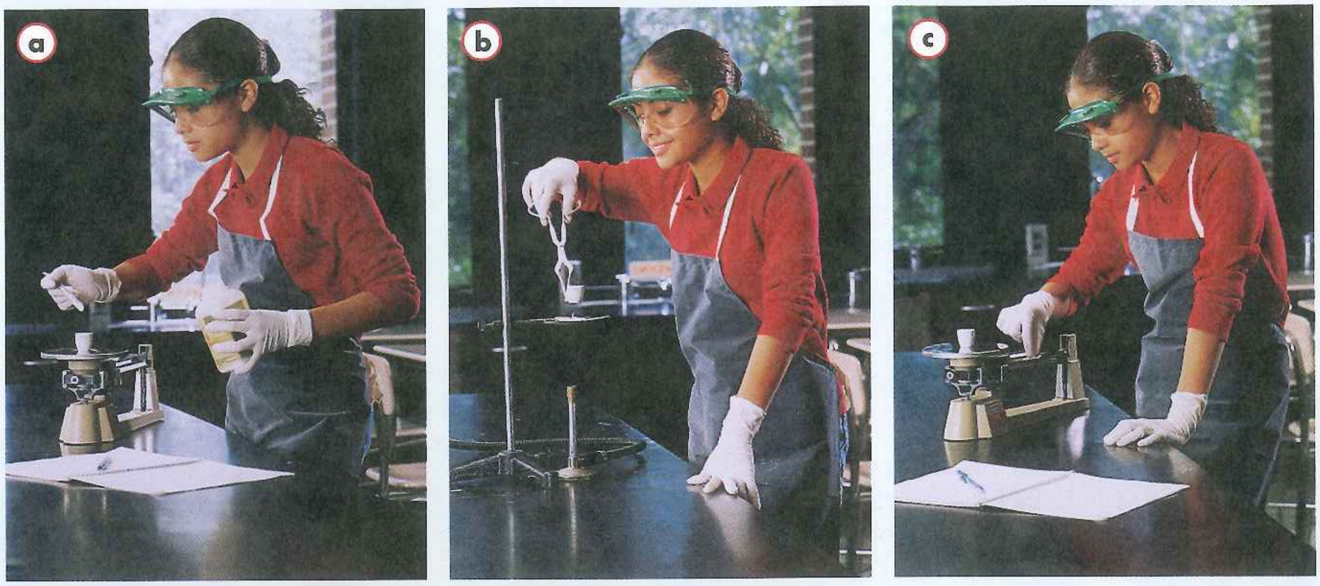


Figure 12.9 Determining Percent Yield
Sodium hydrogen carbonate ($NaHCO_3$) will decompose when heated. **a.** The mass of $NaHCO_3$, the reactant, is measured. **b.** The reactant is heated. **c.** The mass of one of the products, sodium carbonate (Na_2CO_3), the actual yield, is measured. The percent yield is calculated once the actual yield is determined.

Predict *What are the other products of this reaction?*

When a balanced chemical equation is used to calculate the amount of product that will form during a reaction, the calculated value represents the theoretical yield. The **theoretical yield** is the maximum amount of product that could be formed from given amounts of reactants. In contrast, the amount of product that actually forms when the reaction is carried out in the laboratory is called the **actual yield**. The **percent yield** is the ratio of the actual yield to the theoretical yield expressed as a percent.

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

Because the actual yield of a chemical reaction is often less than the theoretical yield, the percent yield is often less than 100 percent. **The percent yield is a measure of the efficiency of a reaction carried out in the laboratory.** This yield is similar to an exam score measuring your efficiency of learning or a batting average measuring your efficiency of hitting a baseball, as in Figure 12.8.

Stoichiometry and conservation of mass dictate that yields of greater than 100 percent are not possible. However, errors and lack of knowledge in a process can cause a reaction to apear to have a yield that is more than 100 percent. For example, if air or water leaks into a system, then more product may be formed than expected.

Many factors can cause percent yields to be less than 100 percent. Reactions do not always go to completion; when a reaction is incomplete, less than the calculated amount of product is formed. Impure reactants and competing side reactions may cause unwanted products to form. Actual yield can also be lower than the theoretical yield due to a loss of product during filtration or in transferring between containers. Moreover, if reactants or products have not been carefully measured, a percent yield of 100 percent is unlikely.

An actual yield is an experimental value. Figure 12.9 shows a typical laboratory procedure for determining the actual yield of a product of a decomposition reaction. For reactions in which percent yields have been determined, you can calculate and, therefore, predict an actual yield if the reaction conditions remain the same.

CHEM TUTOR

Sample Problem 12.10

Calculating the Theoretical Yield of a Reaction

Calcium carbonate, which is found in seashells, is decomposed by heating. The balanced equation for this reaction is

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

What is the theoretical yield of CaO if 24.8 g $CaCO_3$ is heated?

KNOWNS

mass of calcium carbonate = 24.8 g $CaCO_3$

1 mol $CaCO_3$ = 1 mol CaO

UNKNOWN

theoretical yield = ? g CaO

1 Analyze **List the knowns and the unknown.** Calculate the theoretical yield using the mass of the reactant:

g $CaCO_3 \longrightarrow$ mol $CaCO_3 \longrightarrow$ mol CaO $\longrightarrow$ g CaO

2 Calculate **Solve for the unknown.**

Start with the mass of the reactant and convert to moles of the reactant.

$$24.8\ \cancel{g\,CaCO_3} \times \frac{1\ mol\ CaCO_3}{100.1\ \cancel{g\,CaCO_3}}$$

Next, convert to moles of the product using the mole ratio.

$$24.8\ \cancel{g\,CaCO_3} \times \frac{1\ \cancel{mol\,CaCO_3}}{100.1\ \cancel{g\,CaCO_3}} \times \frac{1\ mol\ CaO}{1\ \cancel{mol\,CaCO_3}}$$

Finish by converting from moles to mass of the product.

$$24.8\ \cancel{g\,CaCO_3} \times \frac{1\ \cancel{mol\,CaCO_3}}{100.1\ \cancel{g\,CaCO_3}} \times \frac{1\ \cancel{mol\,CaO}}{1\ \cancel{mol\,CaCO_3}} \times \frac{56.1\ g\ CaO}{1\ \cancel{mol\,CaO}}$$

$$= 13.9\ g\ CaO$$

3 Evaluate **Does the result make sense?** The mole ratio of CaO to $CaCO_3$ is 1:1. The ratio of their masses in the reaction should be the same as the ratio of their molar masses, which is slightly greater than 1:2. The result of the calculations shows that the mass of CaO is slightly greater than half the mass of $CaCO_3$.

30. When 84.8 g of iron(III) oxide reacts with an excess of carbon monoxide, iron is produced.

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

What is the theoretical yield of iron?

If there is an excess of a reactant, then there is more than enough of that reactant and it will not limit the yield of the reaction.

31. When 5.00 g of copper reacts with excess silver nitrate, silver metal and copper(II) nitrate are produced. What is the theoretical yield of silver in this reaction?

Success Stats

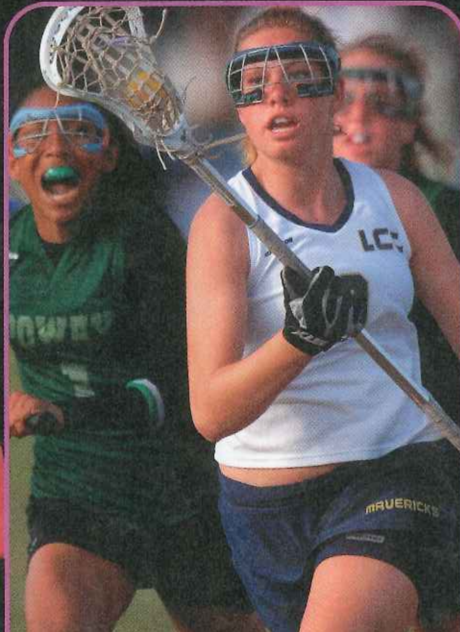
You may not use the term "percent yield" outside of chemistry class, but there are many examples of percent yield in our lives. In chemical reactions, percent yield refers to the amount of product formed in a reaction compared to how much product was possible. In school, percent yield could refer to the graduation rate or a score on a test. In sports, percent yield could refer to the percent of shots that make it into a goal. The actual performance of a product compared to its advertised performance is also an example of percent yield.

Whether it's in the chemistry lab or anywhere else, percent yield is a way to measure how successfully something or someone has performed. The next time you calculate the percent yield of a chemical reaction, think about how this skill could be used in other situations outside of chemistry class.

85 PERCENT PERFORMANCE

Actual Yield: 153 minutes during which a drink stayed hot in thermos

Theoretical Yield: 180 minutes, as advertised by the thermos manufacturer



38 PERCENT SHOT-CONVERSION RATE

Actual Yield: 8 goals scored

Theoretical Yield: 21 shots on goal

Take It Further

1. Calculate Sara's car is advertised to get 43 miles per gallon. Sara calculated her gas mileage over the last month and found that it was 39 miles per gallon. What is the percent yield of Sara's gas mileage.

2. Identify The percent yield of a reaction may be different each time the reaction occurs. Similarly, the performance of an athlete may vary. What are some factors that might affect percent yield?

95 PERCENT GRADUATION RATE

Actual Yield: 305 students graduating

Theoretical Yield: 321 students in the senior class

Sample Problem 12.11

Calculating the Percent Yield of a Reaction

What is the percent yield if 13.1 g CaO is actually produced when 24.8 g $CaCO_3$ is heated?

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

1 Analyze List the knowns and the unknown. Use the equation for percent yield. The theoretical yield for this problem was calculated in Sample Problem 12.10.

KNOWNS

actual yield = 13.1 g CaO

theoretical yield = 13.9 g CaO (from Sample Problem 12.10)

UNKNOWN

percent yield = ? %

2 Calculate Solve for the unknown.

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

Substitute the values for actual yield and theoretical yield into the equation for percent yield.

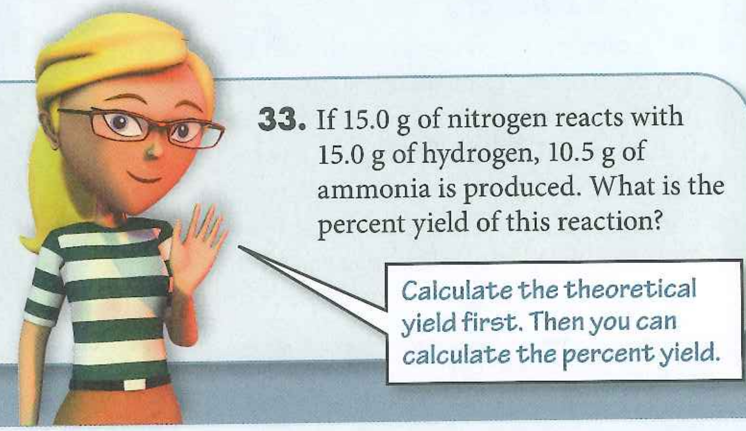
$$\text{percent yield} = \frac{13.1\ \cancel{\text{g CaO}}}{13.9\ \cancel{\text{g CaO}}} \times 100\% = 94.2\%$$

3 Evaluate Does the result make sense? In this example, the actual yield is slightly less than the theoretical yield. Therefore, the percent yield should be slightly less than 100 percent.

32. If 50.0 g of silicon dioxide is heated with an excess of carbon, 27.9 g of silicon carbide is produced.

$$SiO_2(s) + 3C(s) \xrightarrow{\Delta} SiC(s) + 2CO(g)$$

What is the percent yield of this reaction?



33. If 15.0 g of nitrogen reacts with 15.0 g of hydrogen, 10.5 g of ammonia is produced. What is the percent yield of this reaction?

12.3 LessonCheck

34. **Relate Cause and Effect** In a chemical reaction, how does an insufficient quantity of a reactant affect the amount of product formed?

35. **Explain** How can you gauge the efficiency of a reaction carried out in the laboratory?

36. **Define** What is a limiting reagent? An excess reagent?

37. **Calculate** How many grams of SO_3 are produced when 20.0 g FeS_2 reacts with 16.0 g O_2 according to this balanced equation?

$$4FeS_2(s) + 15O_2(g) \longrightarrow 2Fe_2O_3(s) + 8SO_3(g)$$

38. **Calculate** What is the percent yield if 4.65 g of copper is produced when 1.87 g of aluminum reacts with an excess of copper(II) sulfate?

$$2Al(s) + 3CuSO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$$

12 Study Guide

BIGIDEAS

- THE MOLE AND QUANTIFYING MATTER
- REACTIONS

Balanced chemical equations are the basis for stoichiometric calculations. The coefficients of a balanced equation indicate the number of particles, moles, or volumes of gas in the reaction. Mole ratios from the balanced equation are used to calculate the amount of a reactant or product in a chemical reaction from a given amount of one of the reactants or products.

12.1 The Arithmetic of Equations

Chemists use balanced chemical equations as a basis to calculate how much reactant is needed or product is formed in a reaction.

A balanced chemical equation can be interpreted in terms of different quantities, including numbers of atoms, molecules, or moles; mass; and volume.

- stoichiometry (386)

12.2 Chemical Calculations

In chemical calculations, mole ratios are used to convert between a given number of moles of a reactant or product to moles of a different reactant or product.

In a typical stoichiometric problem, the given quantity is first converted to moles. Then, the mole ratio from the balanced equation is used to calculate the moles of the wanted substance. Finally, the moles are converted to any other unit of measurement related to the unit mole.

- mole ratio (390)

Key Equation

mole-mole relationship for $aG \rightarrow bW$:

$$x \text{ mol } G \times \frac{b \text{ mol } W}{a \text{ mol } G} = \frac{xb}{a} \text{ mol } W$$

12.3 Limiting Reagent and Percent Yield

In a chemical reaction, an insufficient quantity of any of the reactants will limit the amount of product that forms.

The percent yield is a measure of the efficiency of a reaction performed in the laboratory.

- limiting reagent (401)
- excess reagent (401)
- theoretical yield (405)
- actual yield (405)
- percent yield (405)

Key Equation

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

Math Tune-Up: Stoichiometry Problems

MATH +/– TUTOR

Problem

Iron metal (Fe) can be obtained from iron ore, Fe_2O_3.

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

How much iron ore is needed to obtain 92.8 grams of iron metal?

1 Analyze

Knowns:
mass of iron = 92.8 g Fe
1 mol Fe_2O_3 / 2 mol Fe (from balanced equation)
1 mol Fe = 55.8 g Fe (molar mass)
1 mol Fe_2O_3 = 159.6 g Fe_2O_3 (molar mass)

Unknown:
Mass of iron ore = ? g Fe_2O_3

2 Calculate

Perform the following steps:
g Fe $\longrightarrow$ mol Fe $\longrightarrow$ mol Fe_2O_3 $\longrightarrow$ g Fe_2O_3

$$92.8\ \cancel{g\ Fe} \times \frac{1\ \cancel{mol\ Fe}}{55.8\ \cancel{g\ Fe}} \times \frac{1\ \cancel{mol\ Fe_2O_3}}{2\ \cancel{mol\ Fe}} \times \frac{159.6\ g\ 1\ mol\ Fe_2O_3}{1\ \cancel{mol\ Fe_2O_3}}$$

$$= 133.0\ g\ Fe_2O_3$$

Hint: Review Sample Problem 12.4 if you have trouble with calculating the mass of a reactant.

3 Evaluate

Since the molar mass of the iron ore is more than twice the molar mass of iron metal, it makes sense that the mass of the iron ore would be greater than the mass of the iron metal produced.

Problem

Sodium hydroxide reacts with carbon dioxide according to the balanced equation below.

$$2NaOH(s) + CO_2(g) \longrightarrow Na_2CO_3(s) + H_2O(l)$$

What is the limiting reagent when 3.50 mol NaOH reacts with 2.00 mol CO_2?

1 Analyze

Knowns:
moles of NaOH = 3.50 mol NaOH
moles of CO_2 = 2.00 mol CO_2

Unknown:
limiting reagent = ?

2 Calculate

Determine how many moles of CO_2 are needed to react with 3.50 mol NaOH.

$$3.50\ mol\ NaOH \times \frac{1\ mol\ CO_2}{2\ mol\ NaOH}$$

$$= 1.75\ mol\ CO_2$$

Only 1.75 mol CO_2 are needed to react with 3.50 mol NaOH. Since there are 2.00 mol CO_2, there is excess CO_2. Therefore, NaOH is the limiting reagent.

3 Evaluate

To check your work, you could start with the given amount of moles of CO_2 and solve for how many moles of NaOH are needed.

Hint: Review Sample Problem 12.8 if you have trouble identifying the limiting reagent.

12 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

12.1 The Arithmetic of Equations

39. Interpret each chemical equation in terms of interacting particles.

a. $2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$
b. $4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(g)$
c. $4K(s) + O_2(g) \longrightarrow 2K_2O(s)$

40. Interpret each equation in Problem 39 in terms of interacting numbers of moles of reactants and products.

41. Calculate and compare the mass of the reactants with the mass of the products for each equation in Problem 39. Show that each balanced equation obeys the law of conservation of mass.

42. Balance the following equation:

$$C_5H_{12}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Interpret the balanced equation in terms of relative number of moles, volumes of gas at STP, and masses of reactants and products.

12.2 Chemical Calculations

43. Explain the term *mole ratio* in your own words. When would you use this term?

44. What ratio is used to carry out each conversion?

a. mol CH_4 to g CH_4
b. L $CH_4(g)$ to mol $CH_4(g)$ (at STP)
c. molecules CH_4 to mol CH_4

★**45.** Carbon disulfide is an important industrial solvent. It is prepared by the reaction of coke with sulfur dioxide.

$$5C(s) + 2SO_2(g) \longrightarrow CS_2(l) + 4CO(g)$$

a. How many moles of CS_2 form when 2.7 mol C reacts?
b. How many moles of carbon are needed to react with 5.44 mol SO_2?
c. How many moles of carbon monoxide form at the same time that 0.246 mol CS_2 forms?
d. How many mol SO_2 are required to make 118 mol CS_2?

★**46.** Methanol (CH_3OH) is used in the production of many chemicals. Methanol is made by reacting carbon monoxide and hydrogen at high temperature and pressure.

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(g)$$

a. How many moles of each reactant are needed to produce 3.60×10^2 g CH_3OH?
b. Calculate the number of grams of each reactant needed to produce 4.00 mol CH_3OH.
c. How many grams of hydrogen are necessary to react with 2.85 mol CO?

47. The reaction of fluorine with ammonia produces dinitrogen tetrafluoride and hydrogen fluoride.

$$5F_2(g) + 2NH_3(g) \longrightarrow N_2F_4(g) + 6HF(g)$$

a. If you have 66.6 g NH_3, how many grams of F_2 are required for a complete reaction?
b. How many grams of NH_3 are required to produce 4.65 g HF?
c. How many grams of N_2F_4 can be produced from 225 g F_2?

48. What information about a chemical reaction is derived from the coefficients in a balanced equation?

49. Rust is produced when iron reacts with oxygen.

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

How many grams of Fe_2O_3 are produced when 12.0 g of iron rusts?

★**50.** Lithium nitride reacts with water to form ammonia and aqueous lithium hydroxide.

$$Li_3N(s) + 3H_2O(l) \longrightarrow NH_3(g) + 3LiOH(aq)$$

a. What mass of water is needed to react with 32.9 g Li_3N?
b. When the above reaction takes place, how many molecules of NH_3 are produced?
c. Calculate the number of grams of Li_3N that must be added to an excess of water to produce 15.0 L NH_3 (at STP).

12.3 Limiting Reagent and Percent Yield

51. What is the significance of the limiting reagent in a reaction? What happens to the amount of any reagent that is present in an excess?

52. How would you identify a limiting reagent in a chemical reaction?

***53.** In a reaction chamber, 3.0 mol of aluminum is mixed with 5.3 mol Cl_2 and reacts. The following balanced chemical equation describes the reaction:

$$2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$$

a. Identify the limiting reagent for the reaction.
b. Calculate the number of moles of product formed.
c. Calculate the number of moles of excess reagent remaining after the reaction.

***54.** Heating an ore of antimony (Sb_2S_3) in the presence of iron gives the element antimony and iron(II) sulfide.

$$Sb_2S_3(s) + 3Fe(s) \longrightarrow 2Sb(s) + 3FeS(s)$$

When 15.0 g Sb_2S_3 reacts with an excess of Fe, 9.84 g Sb is produced. What is the percent yield of this reaction?

55. Phosphoric acid reacts with sodium hydroxide according to the equation:

$$H_3PO_4(aq) + 3NaOH(aq) \longrightarrow Na_3PO_4(aq) + 3H_2O(l)$$

If 1.75 mol H_3PO_4 is made to react with 5.00 mol NaOH, identify the limiting reagent.

Understand Concepts

56. Calcium carbonate reacts with phosphoric acid to produce calcium phosphate, carbon dioxide, and water.

$$3CaCO_3(s) + 2H_3PO_4(aq) \longrightarrow Ca_3(PO_4)_2(aq) + 3CO_2(g) + 3H_2O(l)$$

a. How many grams of phosphoric acid react with excess calcium carbonate to produce 3.74 g $Ca_3(PO_4)_2$?
b. Calculate the number of grams of CO_2 formed when 0.773 g H_2O is produced.

***57.** Nitric acid and zinc react to form zinc nitrate, ammonium nitrate, and water.

$$4Zn(s) + 10HNO_3(aq) \longrightarrow 4Zn(NO_3)_2(aq) + NH_4NO_3(aq) + 3H_2O(l)$$

a. How many atoms of zinc react with 1.49 g HNO_3?
b. Calculate the number of grams of zinc that must react with an excess of HNO_3 to form 29.1 g NH_4NO_3.

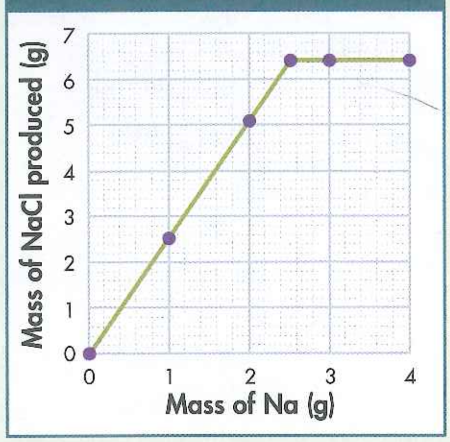
58. If 75.0 g of silderite ore ($FeCO_3$) is heated with an excess of oxygen, 45.0 g of ferric oxide (Fe_2O_3) is produced.

$$4FeCO_3(s) + O_2(g) \longrightarrow 2Fe_2O_3(s) + 4CO_2(g)$$

What is the percent yield of this reaction?

59. In an experiment, varying masses of sodium metal are reacted with a fixed initial mass of chlorine gas. The following graph shows the amounts of sodium used and the amounts of sodium chloride formed.

NaCl Produced by Reacting Sodium With Chlorine



a. Explain the general shape of the graph.
b. Estimate the amount of chlorine gas used in this experiment at the point where the curve becomes horizontal.

***60.** Hydrazine (N_2H_4) is used as rocket fuel. It reacts with oxygen to form nitrogen and water.

$$N_2H_4(l) + O_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

a. How many liters of N_2 (at STP) form when 1.0 kg N_2H_4 reacts with 1.2 kg O_2?
b. How many grams of the excess reagent remain after the reaction?

61. When 50.0 g of silicon dioxide is heated with an excess of carbon, 32.2 g of silicon carbide is produced.

$$SiO_2(s) + 3C(s) \longrightarrow SiC(s) + 2CO(g)$$

a. What is the percent yield of this reaction?
b. How many grams of CO gas are made?

62. If the reaction below proceeds with a 96.8% yield, how many kilograms of $CaSO_4$ are formed when 5.24 kg SO_2 reacts with an excess of $CaCO_3$ and O_2?

$$2CaCO_3(s) + 2SO_2(g) + O_2(g) \longrightarrow 2CaSO_4(s) + 2CO_2(g)$$

63. Ammonium nitrate will decompose explosively at high temperatures to form nitrogen, oxygen, and water vapor.

$$2NH_4NO_3(s) \longrightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$$

What is the total number of liters of gas formed when 228 g NH_4NO_3 is decomposed? (Assume STP.)

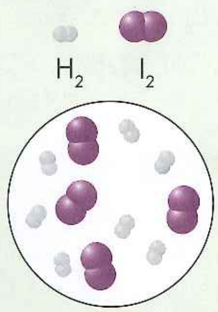
*64. Hydrogen gas can be made by reacting methane (CH_4) with high-temperature steam:

$$CH_4(g) + H_2O(g) \longrightarrow CO(g) + 3H_2(g)$$

How many hydrogen molecules are produced when 158 g of methane reacts with steam?

65. Suppose hydrogen gas and iodine vapor react to give gaseous hydrogen iodide.

a. Write the balanced equation for the reaction.
b. In the atomic window below, which reactant is the limiting reagent?
c. How many molecules of the reagent in excess remain at the completion of the reaction?
d. How many molecules of the limiting reagent need to be added to the atomic window so that all the reactants will react to form products?



66. The following reaction occurs when an automobile battery is charged.

$$PbSO_4(s) + H_2O(l) \longrightarrow PbO_2(s) + Pb(s) + H_2SO_4(aq)$$

a. Balance the equation.
b. How many grams of sulfuric acid are produced when 68.1 g of lead(II) sulfate react?

*67. Liquid sulfur difluoride reacts with fluorine gas to form gaseous sulfur hexafluoride.

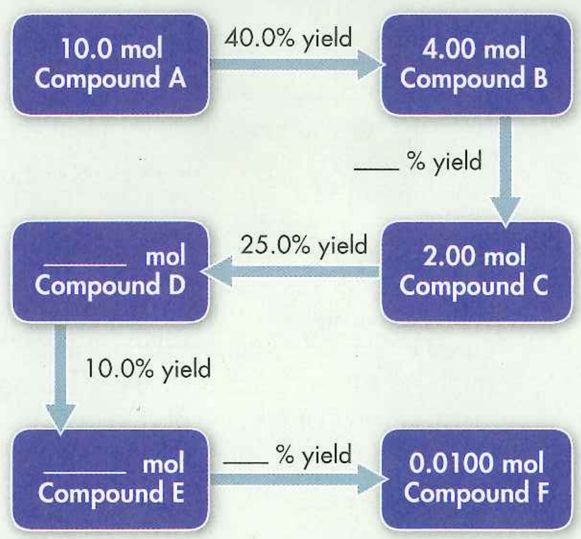
a. Write the balanced equation for the reaction.
b. How many fluorine molecules are required to react with 5.00 mg of sulfur difluoride?
c. What volume of fluorine gas at STP is required to react completely with 6.66 g of sulfur difluoride?

68. Ammonia (NH_3) reacts with oxygen (O_2) to produce nitrogen monoxide (NO) and water.

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$$

How many liters of NO are produced when 1.40 L of oxygen reacts with ammonia?

*69. The manufacture of compound F requires five separate chemical reactions. The initial reactant, compound A, is converted to compound B, compound B is converted to compound C, and so on. The diagram below summarizes the steps in the manufacture of compound F, including the percent yield for each step. Provide the missing quantities or missing percent yields. Assume that the reactant and product in each step react in a one-to-one mole ratio.



Think Critically

70. Evaluate Given a certain quantity of reactant, you calculate that a particular reaction should produce 55 g of a product. When you perform the reaction, you find that you have produced 63 g of product. What is your percent yield? What could have caused a percent yield greater than 100 percent?

71. Explain Would the law of conservation of mass hold in a net ionic equation? Explain.

***72. Calculate** The element phosphorus is manufactured from a mixture of phosphate rock ($Ca_3(PO_4)_2$), sand (SiO_2), and coke (C) in an electric furnace. The chemistry is complex but is summarized by these two equations.

$$Ca_3(PO_4)_2 + SiO_2 \longrightarrow P_4O_{10} + CaSiO_3$$
$$P_4O_{10} + C \longrightarrow P_4 + CO$$

An excess of coke is reacted with 5.5×10^5 g of calcium phosphate and 2.3×10^5 g of sand.

a. Balance each of the equations.
b. What is the limiting reagent?
c. How many grams of phosphorus are produced?
d. How many grams of carbon are consumed?

73. Calculate Sulfuric acid reacts with calcium hydroxide to form calcium sulfate and water.

a. Write the balanced equation for the reaction.
b. Find the mass of unreacted starting material when 75.0 g sulfuric acid reacts with 55.0 g calcium hydroxide.

74. Apply Concepts A car gets 9.2 kilometers to a liter of gasoline. Assuming that gasoline is 100% octane (C_8H_{18}), which has a density of 0.69 g/cm^3, how many liters of air (21% oxygen by volume at STP) will be required to burn the gasoline for a 1250-km trip? Assume complete combustion.

***75. Calculate** Ethyl alcohol (C_2H_5OH) can be produced by the fermentation of glucose ($C_6H_{12}O_6$). If it takes 5.0 h to produce 8.0 kg of alcohol, how many days will it take to consume 1.0×10^3 kg of glucose? (An enzyme is used as a catalyst.)

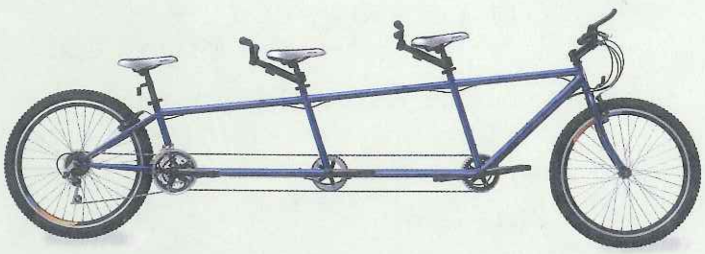
$$C_6H_{12}O_6 \xrightarrow{\text{enzyme}} 2C_2H_5OH + 2CO_2$$

76. Calculate A bicycle built for three has a frame, two wheels, six pedals, and three seats. The balanced equation for this bicycle is

$$F + 2W + 6P + 3S \longrightarrow FW_2P_6S_3$$

How many of each part are needed to make 29 bicycles built for three?

a. frames
b. wheels
c. pedals
d. seats



Enrichment

77. Calculate A 1004.0-g sample of $CaCO_3$ that is 95.0% pure gives 225 L CO_2 at STP when reacted with an excess of hydrochloric acid.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 + H_2O$$

What is the density (in g/L) of the CO_2?

***78. Calculate** The white limestone cliffs of Dover, England, contain a large percentage of calcium carbonate ($CaCO_3$). A sample of limestone with a mass of 84.4 g reacts with an excess of hydrochloric acid to form calcium chloride.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

The mass of calcium chloride formed is 81.8 g. What is the percentage of calcium carbonate in the limestone?

79. Calculate For the reaction below there are 100.0 g of each reactant available. Which reactant is the limiting reagent?

$$2MnO_2 + 4KOH + O_2 + Cl_2 \longrightarrow 2KMnO_4 + 2KCl + 2H_2O$$

80. Calculate The equation for one of the reactions in the process of reducing iron ore to the metal is

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

a. What is the maximum mass of iron, in grams, that can be obtained from 454 g (1.00 lb) of iron(III) oxide?
b. What mass of CO is required to reduce the iron(III) oxide to iron metal?

81. Calculate Esters are a class of compounds that impart a characteristic odor to some fruits. The ester pentyl acetate, composed of carbon, hydrogen, and oxygen, has the odor of bananas. When 7.44 g of this compound undergoes complete combustion, 17.6 g CO_2 and 7.21 g H_2O are produced.

a. What is the empirical formula of pentyl acetate? (*Hint:* All the carbon ends up in the CO_2; all the hydrogen ends up in the H_2O.)
b. The molar mass of pentyl acetate is 130.0 g. What is the molecular formula of this compound?
c. Write the equation for the complete combustion of pentyl acetate.
d. Check your work by using your equation from part c to calculate the grams of CO_2 and H_2O produced by the complete combustion of 7.44 g of pentyl acetate.

*82. **Calculate** Nitric acid, HNO_3, is produced in a complex three-step process summarized by these unbalanced equations.

Step 1: $NH_3 + O_2 \longrightarrow NO + H_2O$

Step 2: $NO + O_2 \longrightarrow NO_2$

Step 3: $NO_2 + H_2O \longrightarrow HNO_3 + NO$

Notice that the nitric oxide, NO, produced in Step 3 is recycled into Step 2.

a. Balance each of the equations.
b. Assuming all the nitrogen from the ammonia will eventually be incorporated into the nitric acid, calculate the mass of nitric acid obtained from 88.0 g NH_3.
c. The concentrated nitric acid used in the lab is a 70.0% by mass solution of HNO_3 in water. Using your answer from part b, calculate the mass of ammonia needed to prepare 1.00 kg of concentrated nitric acid.

83. Calculate SO_3 can be produced in the following two-step process:

$$FeS_2 + O_2 \longrightarrow Fe_2O_3 + SO_2$$

$$SO_2 + O_2 \longrightarrow SO_3$$

Assuming that all the FeS_2 reacts, how many grams of SO_3 are produced when 20.0 g of the FeS_2 reacts with 16.0 g of O_2?

Write About Science

84. Explain Explain this statement: "Mass and atoms are conserved in every chemical reaction, but moles are not necessarily conserved."

85. Explain Review the "mole road map" at the end of Lesson 10.2. Explain how this road map ties into the summary of steps for stoichiometric problems shown in Figure 12.5.

CHEMYSTERY

Cookie Crumbles

Jack tried to make cookies that were extra sweet by adding more sugar than was in the recipe. What Jack didn't realize is that a recipe is like a balanced chemical equation. In order to get the desired product in the reaction of cooking, the reactants, or ingredients, must be combined in specific ratios. Jack changed the amount of sugar, but he didn't change any of the other ingredients. Therefore, he changed the ratios of the ingredients. Balanced chemical equations are important in cooking and in many other fields.

86. Infer If Jack's recipe calls for 2.5 cups of flour and 2 eggs, and Jack wants to scale up the recipe by 50 percent, then how much flour and eggs will he need?

87. Connect to the BIGIDEA How does Jack's baking experience illustrate the concept of a limiting reagent?

Cumulative Review

*88. How many electrons, protons, and neutrons are in an atom of each isotope?

a. titanium-47
b. tin-120
c. oxygen-18
d. magnesium-26

89. When comparing ultraviolet and visible electromagnetic radiation, which has

a. a higher frequency?
b. a higher energy?
c. a shorter wavelength?

90. Identify the larger atom of each pair.

a. sodium and chlorine
b. arsenic and nitrogen
c. fluorine and cesium

91. Write electron dot formulas for the following atoms:

a. Cs
b. Br
c. Ca
d. P

92. Which of these elements form ions with a 2+ charge?

a. potassium
b. sulfur
c. barium
d. magnesium

93. Distinguish among single, double, and triple covalent bonds.

94. Can a compound have both ionic and covalent bonds? Explain your answer.

95. How do you distinguish between a cation and an anion?

96. Name these ions.

a. PO_4^{3-}
b. Al^{3+}
c. Se^{2-}
d. NH_4^+

97. Name each substance.

a. SiO_2
b. K_2SO_4
c. H_2CO_3
d. MgS

98. Write the formula for each compound.

a. aluminum carbonate
b. nitrogen dioxide
c. potassium sulfide
d. manganese(II) chromate
e. sodium bromide

*99. How many grams of beryllium are in 147 g of the mineral beryl ($Be_3Al_2Si_6O_{18}$)?

100. What is the mass, in grams, of a molecule of benzene (C_6H_6)?

*101. What is the molecular formula of oxalic acid, molar mass 90 g/mol? Its percent composition is 26.7% C, 2.2% H, and 71.1% O.

102. How many moles is each of the following?

a. 47.8 g KNO_3
b. 2.22 L SO_2 (at STP)
c. 2.25×10^{22} molecules PCl_3

103. Write a balanced chemical equation for each reaction.

a. When heated, lead(II) nitrate decomposes to form lead(II) oxide, nitrogen dioxide, and molecular oxygen.
b. The complete combustion of isopropyl alcohol (C_3H_7OH) produces carbon dioxide and water vapor.
c. When a mixture of aluminum and iron(II) oxide is heated, metallic iron and aluminum oxide are produced.

104. Balance each equation.

a. $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + NaNO_3(aq)$
b. $AlCl_3(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + Al(NO_3)_3(aq)$
c. $H_2SO_4(aq) + Mg(OH)_2(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$

105. Write a net ionic equation for each reaction in Problem 104.

106. Identify the spectator ions in each reaction in Problem 104.

107. Write a balanced chemical equation for the complete combustion of ribose, $C_5H_{10}O_5$.

If You Have Trouble With . . .

Question	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	104	105	106	107
See Chapter	4	5	6	7	7	8	8	9	9	9	9	10	10	10	10	11	11	11	11	11

Standardized Test Prep

Tips for Success

Anticipate the answer. Use what you know to predict what you think the answer should be. Then look to see if your answer, or one much like it, is given as an option.

Select the choice that best answers each question or completes each statement.

1. Nitric acid is formed by the reaction of nitrogen dioxide with water.

$$3NO_2(g) + H_2O(l) \longrightarrow NO(g) + 2HNO_3(aq)$$

How many moles of water are needed to react with 8.4 mol NO_2?

(A) 2.8 mol (C) 8.4 mol
(B) 3.0 mol (D) 25 mol

2. Phosphorus trifluoride is formed from its elements.

$$P_4(s) + 6F_2(g) \longrightarrow 4PF_3(g)$$

How many grams of fluorine are needed to react with 6.20 g of phosphorus?

(A) 2.85 g (C) 11.4 g
(B) 5.70 g (D) 37.2 g

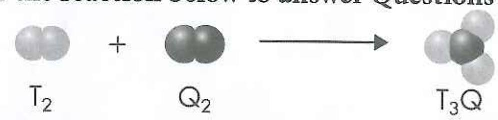
3. Magnesium nitride is formed in the reaction of magnesium metal with nitrogen gas.

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

The reaction of 4.0 mol of nitrogen with 6.0 mol of magnesium produces

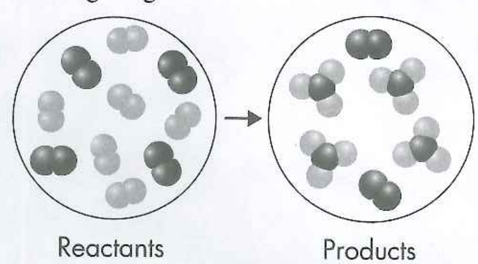
(A) 2.0 mol of Mg_3N_2 and no excess N_2.
(B) 2.0 mol of Mg_3N_2 and 2.0 mol of excess N_2.
(C) 4.0 mol of Mg_3N_2 and 1.0 mol of excess Mg.
(D) 6.0 mol of Mg_3N_2 and 3.0 mol of excess N_2.

Use the reaction below to answer Questions 4 and 5.



4. Write a balanced equation for the reaction between element T and element Q.

5. Based on the atomic windows below, identify the limiting reagent.



For each question, there are two statements. Decide whether each statement is true or false. Then decide whether Statement II is a correct explanation for Statement I.

	Statement I		Statement II
6.	Every stoichiometry calculation uses a balanced equation.	BECAUSE	Every chemical reaction obeys the law of conservation of mass.
7.	A percent yield is always greater than 0% and less than 100%.	BECAUSE	The actual yield in a reaction is never more than the theoretical yield.
8.	The amount of the limiting reagent left after a reaction is zero.	BECAUSE	The limiting reagent is completely used up in a reaction.
9.	The coefficients in a balanced equation represent the relative masses of the reactants and products.	BECAUSE	The mass of the reactants must equal the mass of the products in a chemical reaction.
10.	A mole ratio is always written with the larger number in the numerator.	BECAUSE	A mole ratio will always be greater than 1.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10
See Lesson	12.2	12.2	12.3	12.1	12.3	12.1	12.3	12.3	12.1	12.2