Chapter 9 STOICHIOMETRY

A chemical reaction produces just enough gas to inflate an air bag.

FEATURES

DISCOVER IT! HOW MUCH CAN YOU MAKE?

You need twenty metal paper clips (symbol M), twenty identically colored vinyl-coated paper clips (symbol C), and one plastic sandwich bag.

1. Join together pairs of paper clips of the same color to form models representing 10 diatomic molecules of each reactant. Place these molecules in the plastic bag.
2. Without looking, choose 15 molecules from the plastic bag.
3. Line up the M_2 and C_2 molecules in two adjacent vertical rows.
4. Pair up reactant molecules in the 1 : 3 M_2-to-C_2 ratio as shown in the equation $M_2 + 3C_2 \longrightarrow 2MC_3$.
5. Make the molecules "react" by taking them apart and forming two molecules of the product.
6. Continue making M_2 and C_2 react in a 1:3 ratio until you run out of one of the reactants.

List the number of each type of reactant molecule that was drawn from the bag. How many molecules of the product could you form? Which reactant molecule did you run out of first? How many molecules of each reactant remained at the completion of the reaction? Repeat the experiment and compare the results. Use what you learn in this chapter to provide an explanation for your observations.

THE ARITHMETIC OF EQUATIONS

section 9.1

Silk, one of the most beautiful and luxurious of all fabrics, is spun from the cocoons of tiny silkworms. Silkworms have the unique ability to transform the leaves of mulberry trees into silk thread, which they use to weave their cocoons. More than 3000 cocoons are needed to produce enough silk to make just one elegant Japanese kimono. **How do chemists calculate the amount of reactants and products in chemical reactions?**

objectives

- Calculate the amount of reactants required or product formed in a nonchemical process
- Interpret balanced chemical equations in terms of interacting moles, representative particles, masses, and gas volume at STP

key term

- stoichiometry

Using Everyday Equations

Nearly everything you use is manufactured from chemicals—soaps, shampoos and conditioners, cassette tapes, cosmetics, medicines, and clothes. Obviously, for the manufacturer to make a profit, the cost of making any of these items cannot exceed the money paid for them. Therefore, the chemical processes used in manufacturing must be carried out economically. This is where balanced chemical equations help.

Equations are the recipes that tell chemists what amounts of reactants to mix and what amounts of products to expect. You can determine the quantities of reactants and products in a reaction from the balanced equation. 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. But quantity could just as well be in liters, tons, or molecules. Can you name some other units you might use to measure the amount of matter?

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.

When you bake cookies, you probably use a recipe. A cookie recipe tells you the amounts of ingredients to mix together to make a certain number of cookies. If you need a larger number of cookies than the recipe provides, the amounts of ingredients can be doubled or tripled. In a way, a cookie recipe provides the same kind of information that a balanced chemical equation does. The ingredients are the reactants; the cookies are the products.

Figure 9.1
Just like cooking, manufacturing requires specific amounts of ingredients to get a certain number of products.

Here is another example, this time from the business world rather than from the world of cooking. Imagine you are in charge of manufacturing for the Tiny Tyke Tricycle Company. The business plan for Tiny Tyke requires the production of 128 custom-made tricycles each day. One of your responsibilities is to be sure that there are enough parts available at the start of each day to make these tricycles. 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). The finished tricycle has a "formula" of FSW_3HP_2. The balanced equation for the production of a tricycle is

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

This equation gives you the "recipe" to make a single tricycle: Making a tricycle requires one frame, one seat, three wheels, one handlebar, and two pedals.

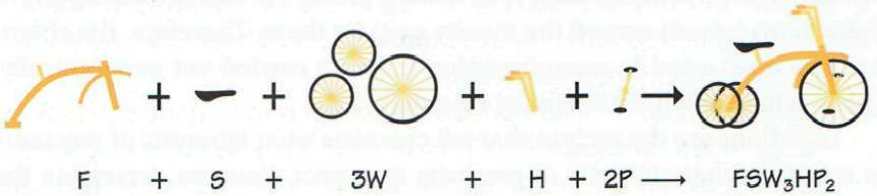


Figure 9.2
A balanced equation can be thought of as a recipe. In the equation shown here, the tricycle parts are the reactants and the assembled tricycle is the product. How many pedals are needed to make four tricycles?

Sample Problem 9-1

In a five-day workweek, Tiny Tyke 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.***

   Knowns:
   - number of tricycles = 640 tricycles
   - 1 FSW_3HP_2 = 3 W (from balanced equation)

   Unknown:
   - number of wheels = ? wheels

   Use the conversion factor $\frac{3\ W}{1\ FSW_3HP_2}$ to calculate the unknown.

2. ***CALCULATE*** ***Solve for the unknown.***

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

3. ***EVALUATE*** ***Does the result make sense?***

   If 3 wheels are required for each tricycle, and a total of 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 cancels with the unit in the denominator of the conversion factor, and the answer is in the unit of the unknown. The conversion factor is exact and does not affect the rounding of the answer.

Practice Problems

1. How many tricycle seats, wheels, and pedals are needed to make 288 tricycles?
2. Write an equation that gives your own "recipe" for making a puppet or a piece of furniture.

Interpreting Chemical Equations

As you may recall from Chapter 7, ammonia is widely used as a fertilizer. Ammonia is produced industrially by the reaction of nitrogen with hydrogen.

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

What kinds of information can be derived from this equation?

1. **Particles** One molecule of nitrogen reacts with three molecules of hydrogen to produce two molecules of ammonia. Nitrogen and hydrogen will always react to form ammonia in this 1:3:2 ratio of molecules. So 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 possible 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 would be the same 1:3 ratio of molecules of reactants. The reaction would form two times Avogadro's number of ammonia molecules.
2. **Moles** You know that Avogadro's number of representative particles is one mole of a substance. On the basis of the particle interpretation you just read, the equation tells you the number of moles of reactants and products. One mole of nitrogen molecules reacts with three moles of hydrogen molecules to form two moles of ammonia molecules. The coefficients of a balanced chemical equation indicate the relative numbers of moles of reactants and products in a chemical reaction. This is the most important information that a balanced chemical equation provides. Using this information, you can calculate the amounts of reactants and products. Does the number of moles of reactants equal the number of moles of product in this reaction?

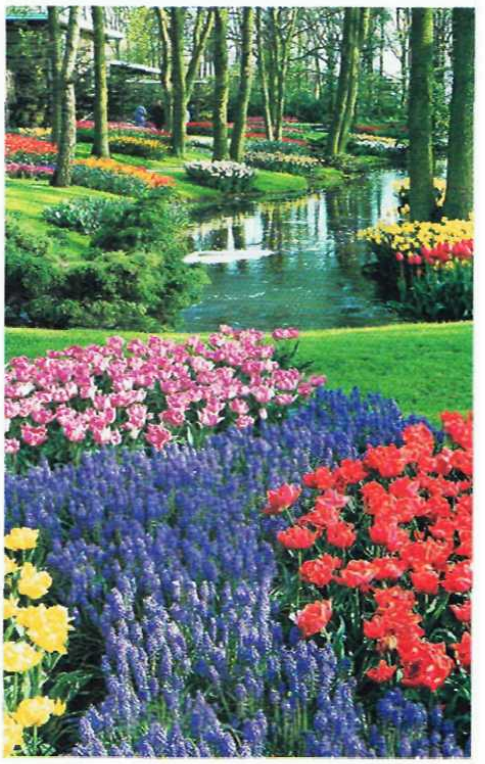


Figure 9.3
Gardeners use ammonium salts as fertilizer. The nitrogen in these salts is essential to plant growth.

$N_2(g)$	+	$3H_2(g)$	→	$2NH_3(g)$
	+		→	
2 atoms N	+	6 atoms H	→	2 atoms N and 6 atoms H
1 molecule N_2	+	3 molecules H_2	→	2 molecules NH_3
10 molecules N_2	+	30 molecules H_2	→	20 molecules NH_3
$1 \times \left(\begin{matrix}6.02 \times 10^{23} \\ \text{molecules } N_2\end{matrix}\right)$	+	$3 \times \left(\begin{matrix}6.02 \times 10^{23} \\ \text{molecules } H_2\end{matrix}\right)$	→	$2 \times \left(\begin{matrix}6.02 \times 10^{23} \\ \text{molecules } NH_3\end{matrix}\right)$
1 mol N_2	+	3 mol H_2	→	2 mol NH_3
28 g N_2	+	3×2 g H_2	→	2×17 g NH_3
		34 g reactants	→	34 g products
Assume STP 22.4 L	+	22.4 L, 22.4 L, 22.4 L	→	22.4 L, 22.4 L
22.4 L N_2		67.2 L H_2		44.8 L NH_3

Figure 9.4
The balanced chemical equation for the formation of ammonia can be interpreted in several ways. How many molecules of NH_3 could be made from 5 molecules N_2 and 15 molecules H_2?

3. **Mass** A balanced chemical equation must obey the law of conservation of mass. This law states that mass can be neither created nor destroyed in an ordinary chemical or physical process. The mole interpretation supports this requirement. Remember that mass is related to the number of atoms in the chemical equation through moles. The mass of 1 mol of nitrogen (28.0 g) plus the mass of 3 mol of hydrogen (6.0 g) does equal the mass of 2 mol of ammonia (34.0 g). So although the number of moles of reactants does not equal the number of moles of product(s), the total number of grams of reactants does equal the total number of grams of product(s). Mass is conserved.
4. **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. It follows that 22.4 L of nitrogen reacts with 67.2 L (3×22.4 L) of hydrogen to form 44.8 L (2×22.4 L) of ammonia.

Look at **Figure 9.4** on the previous page. Do you see that mass and atoms are conserved in this chemical reaction? Mass and atoms are conserved in every chemical reaction. The mass of the reactants equals the mass of the products. The number of atoms of each reactant equals the number of atoms of that reactant in the product(s). Unlike mass and atoms, however, molecules, formula units, moles, and volumes of gases will not necessarily be conserved—although they may. 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 (including the reaction for the formation of ammonia), they are not. Only mass and atoms are conserved in every chemical reaction.

Figure 9.5
Hydrogen sulfide (H_2S) smells like rotten eggs. It escapes from the ground in volcanic areas.

Practice Problems

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

Sample Problem 9-2

Hydrogen sulfide, a foul-smelling gas, is found in nature in volcanic areas. The balanced chemical equation for the burning of hydrogen sulfide is given below. Interpret this equation in terms of the interaction of the following three relative quantities.

a. number of representative particles
b. number of moles
c. masses of reactants and products

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

1. ***ANALYZE*** *Plan a problem-solving strategy.*
   a. The coefficients in the balanced equation give the relative number of molecules of reactants and products.
   b. The coefficients in the balanced equation give the relative number of moles of reactants and products.
   c. A balanced chemical equation obeys the law of conservation of mass. The sum of the masses of the reactants must equal the sum of the masses of the products.

Sample Problem 9-2 (cont.)

2. ***SOLVE*** *Apply the problem-solving strategy.*
   a. 2 molecules H_2S react with 3 molecules O_2 to form 2 molecules SO_2 and 2 molecules H_2O.
   b. 2 mol H_2S react with 3 mol O_2 to produce 2 mol SO_2 and 2 mol H_2O.
   c. Multiply the number of moles of each reactant and product by its molar mass: 2 mol H_2S + 3 mol O_2 $\longrightarrow$ 2 mol SO_2 + 2 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. ***EVALUATE*** *Do the results make sense?*
   Because all the substances in this reaction are molecular, the mole ratio of reactants and products equals the molecular ratio of reactants and products (2:3:2:2). The sum of the masses of the reactants equals the sum of the masses of the products.

Practice Problems (cont.)

4. Balance the equation for the combustion of acetylene:

$$C_2H_2(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

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

Chem ASAP!

Problem-Solving 4

Solve Problem 4 with the help of an interactive guided tutorial.

section review 9.1

5. Your school club has "adopted" a local nursing home and provides welcoming packages to new residents. Each welcoming package contains a toothbrush (B), three washcloths (W), a hand mirror (M), two decks of cards (C), and four small bottles of skin lotion (L).
   a. Write a balanced equation for preparing a welcoming package ($BW_3MC_2L_4$).
   b. Calculate the number of each item needed for 45 packages.
6. Balance this equation: $C_2H_5OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$.
   a. Interpret the equation in terms of numbers of molecules and moles.
   b. Show that the balanced equation obeys the law of conservation of mass.
7. Explain this statement: "Mass and atoms are conserved in every chemical reaction, but moles will not necessarily be conserved."
8. Interpret the following equation in terms of relative numbers of representative particles, numbers of moles, and masses of reactants and products.

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

Chem ASAP! Assessment 9.1 Check your understanding of the important ideas and concepts in Section 9.1.

section 9.2

CHEMICAL CALCULATIONS

Air bags inflate almost instantaneously upon impact. The effectiveness of air bags is based on the rapid conversion of a small mass of sodium azide into a large volume of gas. **How can stoichiometry be used to calculate the volume of gas produced in this reaction?**

objectives

- Construct mole ratios from balanced chemical equations and apply these ratios in mole-mole stoichiometric calculations
- Calculate stoichiometric quantities from balanced chemical equations using units of moles, mass, representative particles, and volumes of gases at STP

Chem ASAP!

Simulation 6

Strengthen your analytical skills by solving stoichiometric problems.

Mole-Mole Calculations

As you just learned, a balanced chemical equation provides a wealth of quantitative information relating representative particles (atoms, molecules, formula units), moles of substances, and masses. Most important, a balanced chemical equation is essential for all calculations involving amounts of reactants and products: If 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.

Figure 9.6

Manufacturing plants produce ammonia by combining nitrogen with hydrogen. Ammonia is used in cleaning products, fertilizers, and in the manufacture of other chemicals.

Look again at the production of ammonia from nitrogen and hydrogen. The balanced equation for the reaction is

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

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. With this interpretation, you can relate moles of reactants to moles of product. The coefficients from the balanced equation are used to write conversion factors called mole ratios. The mole ratios are used to calculate the number of moles of product from a given number of moles of reactant or to calculate the number of moles of reactant from a given number of moles of product. Three of the mole ratios for this equation are

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

What are the other three mole ratios?

CHEMath

DIMENSIONAL ANALYSIS

Converting units is an essential skill for solving problems in chemistry and other sciences. In Section 4.2, you learned how to convert units using *dimensional analysis.* This page provides a chance for you to recall and brush up your skills prior to plunging into Chapter 9—a chapter requiring extensive use of dimensional analysis.

A *conversion factor* is a ratio of equivalent measurements. (Mole ratios, which are used frequently in Chapter 9, are a special type of conversion factor.) By definition, any conversion factor is equal to 1, so you can always multiply by a conversion factor without changing the value of an expression. By carefully choosing conversion factors and canceling units, you can convert any measurement into one expressed in the desired units.

For example, to convert 17 minutes into seconds, you would multiply by a conversion factor that relates the two units. From the relationship 1 minute = 60 seconds, the correct conversion factor can be written and applied.

$$17\ \cancel{\text{min}} \times \frac{60\ \text{s}}{1\ \cancel{\text{min}}} = 1020\ \text{s}$$

Dimensional analysis is also a powerful tool for preventing mistakes. The variables and constants in scientific formulas usually include units, and dimensional analysis can be used to ensure that the units on each side of an equation are the same. For example, consider the formula

$$\text{distance} = \text{velocity} \times \text{time}$$

$$\text{or } d = vt.$$

Notice the typical units in this equation:

$$\text{meters} = \frac{\text{meter}}{\cancel{\text{second}}} \times \cancel{\text{seconds}}$$

After canceling, both sides of the equation have the same units, meters. But if you were to write an incorrect equation using d, v, and t, the units would *not* balance and you would know, immediately, that your formula was incorrect!

Everyone makes mistakes. But if you pay attention to the units in all of your calculations, you will be able to identify and correct many of your mistakes before they cost you a few points and a lower grade.

Example 1

Ingrid drove 18 minutes at 40 km/h. Use $d = vt$ to find out how far she traveled.

Apply the formula: $d = vt = \frac{40\ \text{km}}{1\ \text{h}} \times 18\ \text{min}$

Notice that the time units do not cancel. (Minutes must be converted to hours.) Inserting the proper conversion factor and using dimensional analysis yields,

$$d = vt = \frac{40\ \text{km}}{1\ \cancel{\text{h}}} \times \frac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times 18\ \cancel{\text{min}} = 12\ \text{km}$$

She traveled a distance of 12 km.

Example 2

One mole of copper has a mass of 63.5 g. Find the number of moles in 6.00 kg of copper.

This problem requires two conversion factors. First, kilograms are converted to grams, and then grams are converted to moles.

$$6.00\ \cancel{\text{kg Cu}} \times \frac{1000\ \cancel{\text{g Cu}}}{1\ \cancel{\text{kg Cu}}} \times \frac{1\ \text{mol Cu}}{63.5\ \cancel{\text{g Cu}}} = 94.5\ \text{mol Cu}$$

There are 94.5 moles in 6.00 kg of copper.

Practice Problems

A. Convert a volume of 38.5 L to milliliters.

B. Boris rode a bicycle for 35 minutes at a 24 km/h. How far did he travel?

C. Convert a speed of 90 km/h to meters per second.

D. Convert a speed of 66 ft/s to mi/h. (1 mi = 5280 ft)

E. Convert a density of 158.7 kg/m^3 to g/cm^3.

F. One mole of sulfur has a mass of 32.1 g. Find the mass, in kilograms, of 64.8 moles of sulfur.

G. One mole of 2,2-dichlorohexane ($C_6H_{12}Cl_2$) has a mass of 155.1 g. Find the number of moles in 18.74 kg of 2,2-dichlorohexane.

Sample Problem 9-3

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

1. ***ANALYZE*** ***List the known and the unknown.***

Known:

- moles of nitrogen = 0.60 mol N_2

Unknown:

- moles of ammonia = ? mol NH_3

The conversion is mol $N_2 \rightarrow$ 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.

2. ***CALCULATE*** ***Solve for the unknown.***

$$0.60\ \cancel{\text{mol N}_2} \times \frac{2\ \text{mol NH}_3}{1\ \cancel{\text{mol N}_2}} = 1.2\ \text{mol NH}_3$$

3. ***EVALUATE*** ***Does the result make sense?***

The balanced chemical equation shows that two moles of ammonia are produced for each mole of nitrogen reacted. Note that mole ratios from balanced equations are considered to be exact (defined numbers). They do not enter into the determination of significant figures in the answer.

Practice Problems

9. This equation shows the formation of aluminum oxide.
$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$

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?

10. According to the equation in Problem 9:

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?

Figure 9.7
To determine the number of moles in a sample of a compound, first measure the mass of the sample. Then use the molar mass to calculate the number of moles in that mass.

In the mole ratio below, W is the unknown 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 9-3, is given by

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

Mass-Mass Calculations

No laboratory balance can measure substances directly in moles. Instead, as is shown in **Figure 9.7**, 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, 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.

Sample Problem 9-4

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.*

Knowns:

- mass of hydrogen = 5.40 g H_2
- 3 mol H_2 = 2 mol NH_3 (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

The mass in grams of hydrogen will be used to find the mass in grams of ammonia: g H_2 ⟶ g NH_3.
The coefficients in the balanced equation show that 3 mol H_2 reacts with 1 mol N_2 to produce 2 mol NH_3. The following calculations need to be done:

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

2. ***CALCULATE*** *Solve for the unknown.*

Convert the given (5.40 g H_2) to moles by using the molar mass of hydrogen.

$$5.40\ \cancel{\text{g } H_2} \times \frac{1 \text{ mol } H_2}{2.0\ \cancel{\text{g } H_2}} = 2.7 \text{ mol } H_2$$

Use the mole ratio from the balanced equation to calculate the number of moles of NH_3.

$$2.7\ \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3\ \cancel{\text{mol } H_2}} = 1.8 \text{ mol } NH_3$$

Convert the moles of NH_3 to grams of NH_3 by using the molar mass of ammonia.

$$1.8\ \cancel{\text{mol } NH_3} \times \frac{17.0 \text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} = 31 \text{ g } NH_3$$

This series of calculations can be combined:

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

$$\underset{\text{Given quantity}}{5.40\ \cancel{\text{g } H_2}} \times \underset{\text{Change given unit to moles}}{\frac{1\ \cancel{\text{mol } H_2}}{2.0\ \cancel{\text{g } H_2}}} \times \underset{\text{Mole ratio}}{\frac{2\ \cancel{\text{mol } NH_3}}{3\ \cancel{\text{mol } H_2}}} \times \underset{\text{Change moles to grams}}{\frac{17.0 \text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}}} = 31 \text{ g } NH_3$$

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.

Practice Problems

11. 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?

12. Using the same equation, determine how many moles of CaC_2 are needed to react completely with 49.0 g H_2O.

Chem ASAP!

Problem-Solving 11

Solve Problem 11 with the help of an interactive guided tutorial.

Figure 9.8

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

If the law of conservation of mass is true, how is it possible to make 31 g NH_3 from only 5.40 g H_2? Looking back at the equation for the reaction, you will see that hydrogen is not the only reactant. Another reactant, nitrogen, is also involved. If you were to calculate the number of grams of nitrogen needed to produce 31 g NH_3 and then compare the total masses of reactants and products, you would have an answer to this question. Go ahead and try it!

You can see from Sample Problem 9-4 that mass-mass problems can be solved in basically the same way as mole-mole problems. **Figure 9.9** reviews the steps for the mass-mass conversion of any given mass *(G)* and any wanted mass *(W)*.

1. The mass G is changed 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. The moles of *G* are changed 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. The moles of W are changed 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 9.9 also shows 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 9.9** as they are appropriate to the problem you are solving.

LINK TO AGRICULTURE

Ammonia in the Nitrogen Cycle

Ammonia is part of the nitrogen cycle in nature. Earth's atmosphere contains 0.01 parts per million of ammonia, and small

amounts of ammonia occur in volcanic gases. Most ammonia cycles through the living world without returning to the atmosphere. Ammonia plays a role in several stages of the nitrogen cycle. Nitrogen-fixing bacteria form nodules, or swellings, on the roots of plants in the legume family, such as beans and clover plants. Here the bacteria change atmospheric nitrogen into ammonia molecules or ammonium ions. Other bacteria break down the nitrogenous material in dead plants and animals into ammonia molecules. Certain soil bacteria oxidize these molecules into nitrate ions, the form readily absorbed by plant roots. When a plant dies, this cycle begins again.

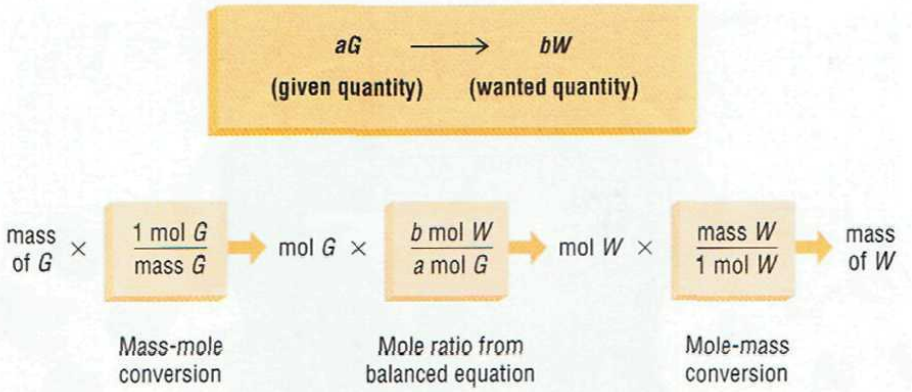


Figure 9.9
This general solution diagram indicates the steps necessary to solve a mass-mass stoichiometry problem: convert mass to moles, use the mole ratio, and then convert moles to mass. Is the given always a reactant?

Other Stoichiometric Calculations

As you already know, a balanced chemical equation indicates the relative number of moles of reactants and products. From this foundation, stoichiometric calculations can be expanded to include any unit of measurement 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, volume-volume, and particle-mass calculations. In any of these problems, 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. Once this has been determined, the moles are converted to any other unit of measurement related to the unit mole, as the problem requires. **Figure 9.10** summarizes these steps for a typical stoichiometric problem.

Figure 9.10
With your knowledge of conversion factors and this problem-solving approach you can solve a variety of stoichiometric problems. What conversion factor is used to convert moles to representative particles?

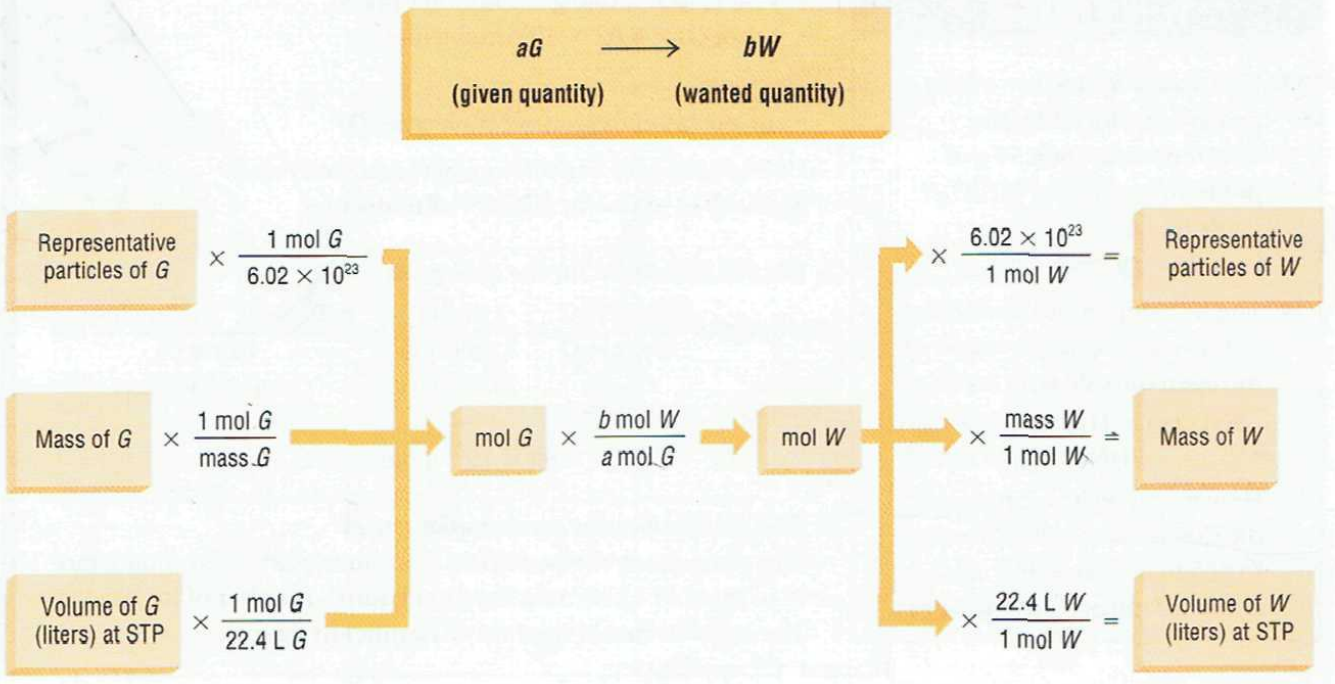
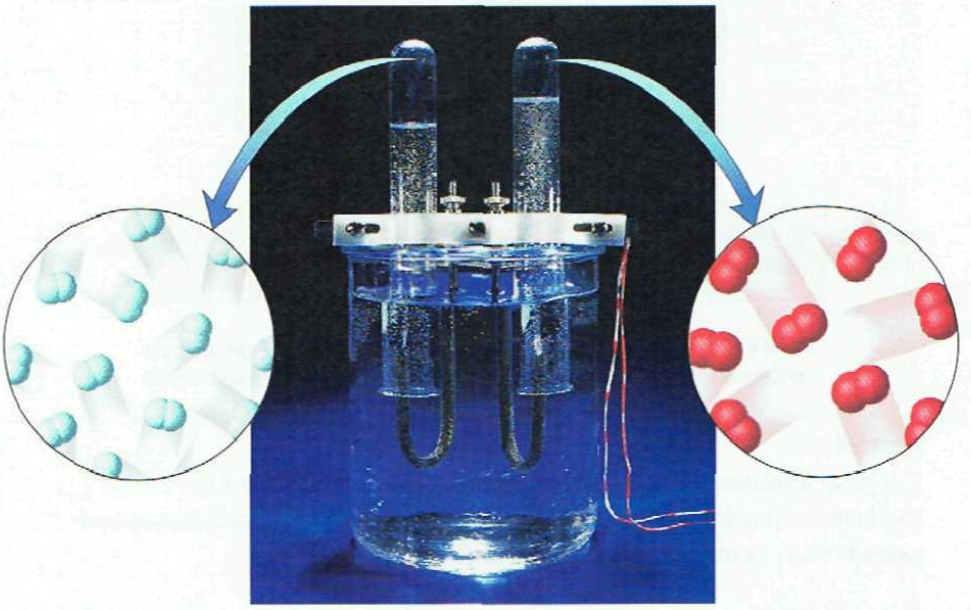


Figure 9.11
The electrolysis of water causes it to decompose into hydrogen and oxygen.



Sample Problem 9-5

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

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

1. *ANALYZE* *List the knowns and the unknown.*

Knowns:

- mass of water = 29.2 g H_2O
- 2 mol H_2O = 1 mol O_2 (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 = ? molecules O_2

Use appropriate conversion factors to convert the given quantity in grams to number of molecules.

2. *CALCULATE* *Solve for the unknown.*

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

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

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.

Practice Problems

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

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

14. 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 9-6

Assuming STP, how many liters of oxygen are needed to produce 19.8 L 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.***

Knowns:

- volume of sulfur trioxide = 19.8 L
- 2 mol SO_3 = 1 mol O_2 (from balanced equation)
- 1 mol SO_3 = 22.4 L SO_3 (at STP)
- 1 mol O_2 = 22.4 L O_2 (at STP)

Unknown:

- volume of oxygen = ? L O_2

2. ***CALCULATE*** ***Solve for the unknown.***

$$\underset{\text{Given quantity}}{19.8\ \cancel{\text{L SO}_3}} \times \underset{\text{Change to moles}}{\frac{1\ \cancel{\text{mol SO}_3}}{22.4\ \cancel{\text{L SO}_3}}} \times \underset{\text{Mole ratio}}{\frac{1\ \cancel{\text{mol O}_2}}{2\ \cancel{\text{mol SO}_3}}} \times \underset{\text{Change to liters}}{\frac{22.4\ \text{L O}_2}{1\ \cancel{\text{mol O}_2}}} = 9.90\ \text{L O}_2$$

3. ***EVALUATE*** ***Does the result make sense?***

Because 2 mol SO_3 is produced for each 1 mol O_2 that reacts, the volume of O_2 should be half the volume of SO_3. The answer should have three significant figures.

Practice Problems

15. 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?

16. 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?

Chem ASAP!

Problem-Solving 16

Solve Problem 16 with the help of an interactive guided tutorial.

In Sample Problem 9-6, did you notice that the 22.4 L/mol factors canceled out? This will always be true in a volume-volume problem. The coefficients in a balanced chemical equation indicate the relative numbers of moles. The coefficients also indicate the relative volumes of interacting gases. The volume can be expressed in any unit. What are some other units of volume?

Sample Problem 9-7

Nitrogen monoxide and oxygen gas combine to form the brown gas nitrogen dioxide. How many milliliters of nitrogen dioxide are produced when 3.4 mL of oxygen reacts with an excess of nitrogen monoxide? Assume conditions of STP.

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

1. ***ANALYZE*** ***List the knowns and the unknown.***

Knowns:

- volume of oxygen = 3.4 mL O_2
- 1 mL O_2 = 2 mL NO_2 (from balanced equation)

Unknown:

- volume of nitrogen dioxide = ? mL NO_2

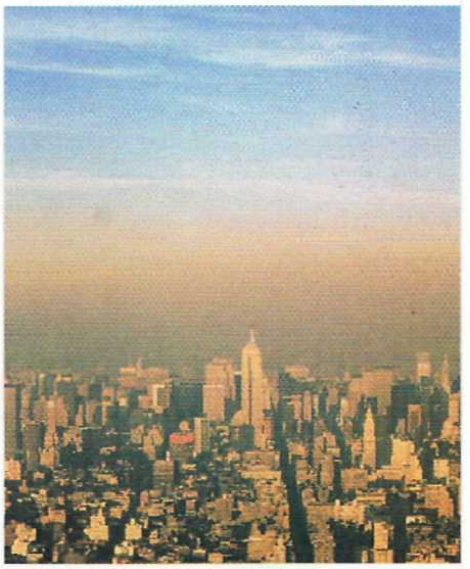


Figure 9.12

The brown gas nitrogen dioxide is a component of photochemical smog, which builds up when still air hangs over a city or other pollution source. Persistent smog can pose a health danger.

Sample Problem 9-7 (cont.)

2. ***CALCULATE*** ***Solve for the unknown.***

$$3.4\ \cancel{\text{mL } O_2} \times \frac{2\ \text{mL } NO_2}{1\ \cancel{\text{mL } O_2}} = 6.8\ \text{mL } NO_2$$

Given quantity — Volume ratio

3. ***EVALUATE*** ***Does the result make sense?***
The conversion is mL $O_2 \longrightarrow$ mL NO_2. The given quantity of O_2 is multiplied by the volume ratio from the balanced equation that allows the given unit to cancel. Because the volume ratio is 2 volumes NO_2 to 1 volume O_2, the calculated volume of NO_2 should be twice the given volume of O_2. The answer should have two significant figures.

Practice Problems

Consider this equation.

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

17. Calculate the volume of sulfur dioxide produced when 27.9 mL O_2 reacts with carbon disulfide.
18. How many deciliters of carbon dioxide are produced when 0.38 L SO_2 is formed?

section review 9.2

19. Isopropyl alcohol (C_3H_7OH) burns in air according to this equation:

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

   a. Calculate the moles of oxygen needed to react with 3.40 mol C_3H_7OH.
   b. Find the moles of each product formed when 3.40 mol C_3H_7OH reacts with oxygen.

20. 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

21. 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 burns?
   b. How many grams of oxygen are required to burn 52.0 g C_2H_2?
   c. Use the answers from **a** and **b** to show that this equation obeys the law of conservation of mass.

22. Tin(II) fluoride, formerly found in many kinds of toothpaste, is formed in this reaction:

$$Sn(s) + 2HF(g) \longrightarrow SnF_2(s) + H_2(g)$$

   a. How many liters of HF are needed to produce 9.40 L H_2 at STP?
   b. How many molecules of H_2 are produced by the reaction of tin with 20.0 L HF at STP?
   c. How many grams of SnF_2 can be made by reacting 7.42×10^{24} molecules of HF with tin?

portfolio project

Using the Internet or the library, research analytical balances. How small a mass or large a mass can these types of balances measure? Compare different balances to find out whether there is a relationship between precision and cost.

Chem ASAP! Assessment 9.2 Check your understanding of the important ideas and concepts in Section 9.2.

SMALL-SCALE LAB

ANALYSIS OF BAKING SODA

SAFETY

Wear safety glasses and follow the standard safety procedure as outlined on page 18.

PURPOSE

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

MATERIALS

- baking soda
- 3 plastic cups
- pipets of HCl, NaOH, and thymol blue
- soda straw
- balance
- pH sensor (optional)

PROCEDURE

Sensor version available in the Probeware Lab Manual.

Prepare and analyze a sample of baking soda. Take care to write down the results of each step.

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 ¾ 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.

ANALYSIS

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

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

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

   (Step B − Step A)

3. 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 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 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 the mass of the baking soda from the reaction data.

   (0.084 g/mmol × Step 5)

7. Calculate the percent error of the experiment.

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

YOU'RE THE CHEMIST

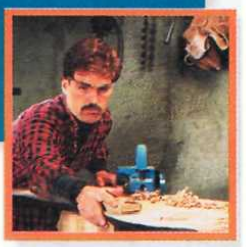
The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Analyze It!** *For each calculation you did, substi*tute each quantity (number and unit) into the equation and cancel the units to explain why each step gives the quantity desired.

2. **Design It!** 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.

section 9.3

LIMITING REAGENT AND PERCENT YIELD



objectives

- Identify and use the limiting reagent in a reaction to calculate the maximum amount of product(s) produced and the amount of excess reagent
- Calculate theoretical yield, actual yield, or percent yield given appropriate information

key terms

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

If a carpenter had two table tops 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 the limiting factor in the construction of four-legged tables. A similar concept applies in chemistry when knowing the exact amounts of reactants and products in a chemical reaction is crucial. **How does a limiting reagent affect a chemical reaction?**

What Is a Limiting Reagent?

Perhaps you know from your own experience that many cooks follow a recipe when making a new dish. They know that sufficient quantities of all the ingredients must be available. Suppose, for example, that you are preparing to make lasagna and you have more than enough meat, tomato sauce, ricotta cheese, eggs, mozzarella cheese, spinach, and seasoning on hand. However, you have only half a box of lasagna noodles. The amount of lasagna you can make will be limited by the amount of noodles you have. Thus the noodles are the limiting reagent in this baking venture. **Figure 9.13** illustrates another example of a limiting reagent in the kitchen. A chemist often faces a similar situation. It is impossible for a chemist to make a certain amount of a desired compound if there is an insufficient quantity of any of the required reactants.

As you know, a balanced chemical equation is a chemist's recipe—a recipe that can be interpreted on a microscopic scale (interacting particles) or on a macroscopic scale (interacting moles). The coefficients used to write

Figure 9.13

The amount of product is determined by the quantity of the limiting reagent. In this example, the rolls are the limiting reagent. No matter how much of the other ingredients you have, with two rolls you can make only two sandwiches.

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 9.14** 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. It is called the limiting reagent. As the name implies, the **limiting reagent** limits or determines the amount of product that can be formed in 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 will remain unreacted. You probably know that if you put a glass over a burning candle, the candle goes out. In this example of the combustion of candle wax, what is the limiting reagent?

Sometimes in 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 each reactant to moles. Then the limiting reagent can be identified, as in Sample Problem 9-8. Finally, the amount of product can be determined from the given amount of limiting reagent.

Chem ASAP!

Animation 8

Apply the limiting reagent concept to the production of iron from iron ore.

Chemical Equations

	$N_2(g)$	+	$3H_2(g)$	→	$2NH_3(g)$
"Microscopic recipe"	1 molecule N_2	+	3 molecules H_2	→	2 molecules NH_3
"Macroscopic recipe"	1 mol N_2	+	3 mol H_2	→	2 mol NH_3

Experimental Conditions

	Reactants		Products
Before reaction	2 molecules N_2	3 molecules H_2	0 molecules NH_3
After reaction	1 molecule N_2	0 molecules H_2	2 molecules NH_3

Figure 9.14

The "recipe" calls for 3 molecules of H_2 for every 1 molecule of N_2. In this particular experiment, H_2 is the limiting reagent and N_2 is in excess. Would the amount of products formed change if you started with 4 molecules of N_2 and 3 molecules of H_2?

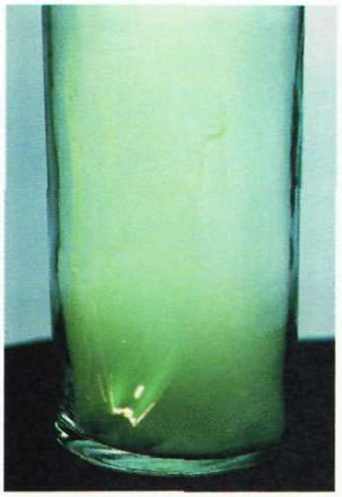


Figure 9.15
Igniting sodium metal in chlorine gas produces a white smoke of NaCl and bright yellow light.

Sample Problem 9-8

Sodium chloride can be prepared by the reaction of sodium metal with chlorine gas.

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Suppose that 6.70 mol Na reacts with 3.20 mol Cl_2.

a. What is the limiting reagent?

b. How many moles of NaCl are produced?

1. ANALYZE ***List the knowns and the unknown for* a.**

Knowns:

- moles sodium = 6.70 mol Na
- moles chlorine = 3.20 mol Cl_2
- 2 mol Na = 1 mol Cl_2 (from balanced equation)

Unknown:

- limiting reagent = ?

The known amount of one of the reactants is multiplied by the mole ratio from the balanced equation to calculate the required amount of the other reactant. Sodium is chosen arbitrarily here: mol Na $\longrightarrow$ mol Cl_2.

2. CALCULATE ***Solve for the unknown.***

$$\underbrace{6.70\ \cancel{\text{mol Na}}}_{\text{Given amount}} \times \underbrace{\frac{1\ \text{mol Cl}_2}{2\ \cancel{\text{mol Na}}}}_{\text{Mole ratio}} = \underbrace{3.35\ \text{mol Cl}_2}_{\text{Required amount}}$$

This calculation indicates that 3.35 mol Cl_2 is needed to react with 6.70 mol Na. Because only 3.20 mol Cl_2 is available, however, chlorine becomes the limiting reagent. Sodium, then, must be in excess.

1. ANALYZE ***List the knowns and the unknown for* b.**

Knowns:

- amount of limiting reagent = 3.20 mol Cl_2
- 1 mol Cl_2 = 2 mol NaCl (from balanced equation)

Unknown:

- moles of sodium chloride = ? mol NaCl

2. CALCULATE ***Solve for the unknown.***

$$3.20\ \cancel{\text{mol Cl}_2} \times \frac{2\ \text{mol NaCl}}{1\ \cancel{\text{mol Cl}_2}} = 6.40\ \text{mol NaCl}$$

3. EVALUATE ***Do the results make sense?***

Because the ratio of the given moles of sodium to chlorine was greater than 2:1, which is the ratio from the balanced equation, sodium should be in excess and chlorine should be the limiting reagent.

Practice Problems

23. 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 is reacted with 6.30 mol O_2,

a. identify the limiting reagent.

b. calculate the moles of water produced.

24. The equation for the incomplete combustion of ethene (C_2H_4) is

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

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.

Chem ASAP!

Problem-Solving 24
Solve Problem 24 with the help of an interactive guided tutorial.

Sample Problem 9-9

As illustrated in **Figure 9.16**, the properties of copper(I) sulfide are very different from the properties of the elements copper and sulfur.

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

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

b. What is the maximum number of grams of Cu_2S that can be formed?

1. ANALYZE ***List the knowns and the unknown for* a.**

Knowns:

- mass of copper = 80.0 g Cu
- mass of sulfur = 25.0 g S
- 2 mol Cu = 1 mol S (from balanced equation)

Unknown:

- limiting reagent = ?

The number of moles of each reactant must first be found:

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

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:

$$\text{mol Cu} \longrightarrow \text{mol S}$$

2. CALCULATE ***Solve for the unknown.***

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

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

$$1.26\ \cancel{\text{mol Cu}} \times \frac{1\ \text{mol S}}{2\ \cancel{\text{mol Cu}}} = 0.630\ \text{mol S}$$

Given quantity — Mole ratio — Needed amount

Comparing the amount of sulfur needed (0.630 mol S) with the given amount (0.779 mol S) indicates that sulfur is in excess. Thus copper is the limiting reagent.

1. ANALYZE ***List the knowns and the unknown for* b.**

Knowns:

- limiting reagent = 1.26 mol Cu
- 2 mol Cu = 1 mol Cu_2S (from balanced equation)
- 1 mol Cu_2S = 159.1 g Cu_2S (molar mass)

Unknown:

- mass copper(I) sulfide = ? g Cu_2S

The limiting reagent, which was determined in the last step, is used to calculate the maximum amount of Cu_2S formed:

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

Figure 9.16
Black crystalline copper(I) sulfide (bottom) is formed as a product when the reactants, sulfur (top left) and copper (top right), are heated together.

Practice Problems

25. Hydrogen gas can be produced in the laboratory by the reaction of magnesium metal with hydrochloric acid.

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

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

b. How many grams of hydrogen can be produced when 6.00 g HCl is added to 5.00 g Mg?

Sample Problem 9-9 (cont.)

2. ***CALCULATE*** *Solve for the unknown.*

$$1.26\ \text{mol Cu} \times \frac{1\ \text{mol Cu}_2\text{S}}{2\ \text{mol Cu}} \times \frac{159.1\ \text{g Cu}_2\text{S}}{1\ \text{mol Cu}_2\text{S}} = 1.00 \times 10^2\ \text{g Cu}_2\text{S}$$

The given quantity of copper, 80.0 g, could have been used for this step instead of the moles of copper, which were calculated in the very first step of the solution.

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. The amount of Cu_2S produced should be less than the sum of copper and sulfur (80.0 g Cu + 25.0 g S = 105.0 g) because sulfur was in excess.

Practice Problems (cont.)

26. Acetylene (C_2H_2) will burn in the presence of 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 7.4 mol O_2?

Calculating the Percent Yield

In theory, when a teacher gives an exam to the class, every student should get a grade of 100%. For a variety of reasons, this 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 quantity that shows how well you did on the exam (questions answered correctly) compared with how well you could have done if you had answered all the questions correctly (100%). This calculation is analogous to the percent yield calculation that you do in the laboratory when the product from a chemical reaction is less than you expected based on the balanced chemical equation.

So far in this chapter, you have probably assumed that when doing stoichiometric problems things do not go wrong in chemical reactions. This assumption is as faulty as assuming that all students will score 100% on an exam. When an equation is used to calculate the amount of product that will form during a reaction, a value representing the theoretical yield is obtained. 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 actual yield is often less than the theoretical yield. The **percent yield** is the ratio of the actual yield to the theoretical yield expressed as a percent. The percent yield measures the efficiency of the reaction.

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

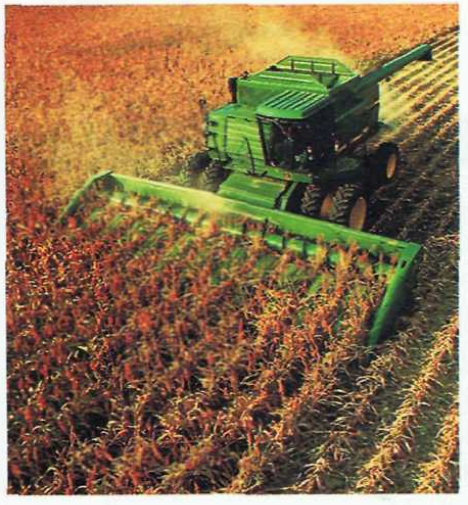


Figure 9.17
The productivity of a farm is measured in yield. Because growing conditions may vary from year to year, the actual yield often differs from the theoretical yield.

A percent yield should not normally be larger than 100%. Many factors cause percent yields to be less than 100%. Reactions do not always go to completion; when this occurs, 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

Figure 9.18
A batting average is actually a percent yield.

between containers. Moreover, if reactants or products have not been carefully measured, a percent yield of 100% is unlikely.

An actual yield is an experimental value. **Figure 9.19** shows a typical laboratory procedure for determining the actual yield of a product of a decomposition reaction. If you do not do an experiment, you cannot calculate a percent yield unless you are given the value of an actual yield. 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. A farmer's crop yield could also be expressed as a percent yield. What factors would a farmer use to predict a theoretical yield?

Figure 9.19
(a) *The mass of sodium hydrogen carbonate ($NaHCO_3$), 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 after the reaction is completed. The percent yield can be calculated once the actual yield has been determined. What are the other products of this reaction?*

(a) (b) (c)

Sample Problem 9-10

Calcium carbonate is decomposed by heating, as shown in the following equation.

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

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

b. What is the percent yield if 13.1 g CaO is produced?

1. *ANALYZE* ***List the knowns and the unknown for*** **a.**

Knowns:

- mass of calcium carbonate = 24.8 g $CaCO_3$
- 1 mol $CaCO_3$ = 1 mol CaO (from balanced equation)
- 1 mol $CaCO_3$ = 100.1 g $CaCO_3$ (molar mass)
- 1 mol CaO = 56.1 g CaO (molar mass)

Unknown:

- theoretical yield of calcium oxide = ? g CaO

The theoretical yield can be calculated using the mass of the reactant:

$$\text{g } CaCO_3 \longrightarrow \text{mol } CaCO_3 \longrightarrow \text{mol CaO} \longrightarrow \text{g CaO}$$

2. *CALCULATE* ***Solve for the unknown.***

First, the theoretical yield of the reaction is calculated.

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

$$= 13.9 \text{ g CaO}$$

1. *ANALYZE* ***List the knowns and the unknown for*** **b.**

Knowns:

- actual yield = 13.1 g CaO
- theoretical yield = 13.9 g CaO (from **a**)
- $\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$

Unknown:

- percent yield = ? %

2. *CALCULATE* ***Solve for the unknown.***

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

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

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%. The answer should have three significant figures.

Practice Problems

27. When 84.8 g of iron(III) oxide reacts with an excess of carbon monoxide, 54.3 g of iron is produced.

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

What is the percent yield of this reaction?

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

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

What is the percent yield of this reaction?

Chem ASAP!

Problem-Solving 28

Solve Problem 28 with the help of an interactive guided tutorial.

MINI LAB

Limiting Reagents

PURPOSE

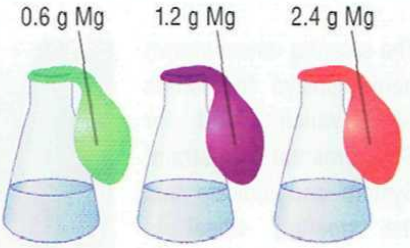
To illustrate the concept of a limiting reagent in a chemical reaction.

MATERIALS

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

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 metal down into each flask. Observe the volume of gas produced until the reaction in each flask is completed.



ANALYSIS AND CONCLUSIONS

1. 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. Write a balanced equation for the reaction between magnesium metal and hydrochloric acid.
3. 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. 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?

section review 9.3

29. What is a limiting reagent? An excess reagent?

30. 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?

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

31. What is the difference between an actual yield and a theoretical yield? Which yield is larger for a given reaction? How are these values used to determine percent yield?

32. 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?

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

Chem ASAP! Assessment 9.3 Check your understanding of the important ideas and concepts in Section 9.3.

Chemistry Serving...Society

JUST THE RIGHT VOLUME OF GAS

The steering wheel shown here displays the letters SRS, which stand for Supplemental Restraint System and indicate that the steering wheel is equipped with an air bag. If you are lucky, you will never have to experience an air bag. But if you do have a front-end collision while driving, an air bag may save your life. A sensor in the front of the car detects the sudden deceleration and sends a signal to a cylinder containing a mixture of chemicals. In the cylinder, an igniter goes off, starting a series of chemical reactions that release a large volume of nitrogen gas. The gas fills the air bag, and you hit the soft bag instead of the steering wheel or dashboard. All this happens in less than a second.

The design of the air bag may sound fairly simple, but it was actually quite a challenge. Engineers had to come up with a system that was safe and effective and that would still work after many years of not being used. Perhaps most importantly, the bag had to inflate in less than a tenth of a second, and it had to inflate with exactly the right amount of gas. If it underinflated it would not provide enough protection; if it overinflated it would cause injury or it might even rupture. To design an air bag inflation system that met these requirements, engineers had to choose the right chemical reactants and pay close attention to the stoichiometry of the reactions.

The engineers had to determine the exact quantity of each reactant to include in the reaction mixture and to allow for the change in gas density caused by the heat released in the reaction.

The inflation system that engineers settled on uses a series of chemical reactions that take place very quickly. The first reaction—set off by the igniter—is the decomposition of sodium azide into sodium metal and nitrogen gas

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

By itself, this reaction cannot fill the air bag fast enough, and the sodium metal that is produced is dangerously reactive. To solve these problems, the engineers included potassium nitrate in the mixture of reactants. The potassium nitrate reacts with the sodium produced in the first reaction, releasing even more nitrogen gas

$$10Na + 2KNO_3 \longrightarrow K_2O + 5Na_2O + N_2$$

The heat released by this reaction raises the temperature of the gaseous product, helping the bag inflate even faster. The heat causes all the solid reaction products to fuse together with SiO_2, powdered sand, which is also part of the reaction mixture. This forms a safe, unreactive glass.

The volume of gas produced by these reactions is the key factor in getting the proper bag inflation. The volume depends on both the quantity of reactants and the temperature of the gas. Thus, the engineers had to determine the exact quantity of each reactant to include in the reaction mixture and to allow for the change in gas pressure caused by the heat released in the reaction. Even the quantity of potassium nitrate included in the mixture had to be precisely determined to make sure no sodium metal was left over from the first reaction. So if an air bag ever saves you from injury in a front-end collision, be thankful that the chemical engineers who designed the air bag's inflation system knew their stoichiometry!

CHEMISTRY IN CAREERS

QUALITY CONTROL CHEMIST
Interested in a career maintaining high-quality standards?
See page 872.

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CHEMICAL SPECIALIST
Local food service distributor see responsible self-motivated indivi

CHEMIST NEEDED

Chapter 9 STUDENT STUDY GUIDE

Take It to the NET
For interactive study and review, go to www.phschool.com

KEY TERMS

- actual yield *p. 256*
- excess reagent *p. 253*
- limiting reagent *p. 253*
- percent yield *p. 256*
- stoichiometry *p. 237*
- theoretical yield *p. 256*

KEY EQUATIONS AND RELATIONSHIPS

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

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

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

CONCEPT SUMMARY

9.1 The Arithmetic of Equations

- The coefficients in a balanced chemical equation tell the relative number of moles of reactants and products.
- Chemists use moles to do chemical arithmetic, or stoichiometry.
- All stoichiometric calculations involving chemical reactions begin with a balanced equation because mass is conserved in every chemical reaction.
- The number and kinds of atoms in the reactants equal the number and kinds of atoms in the products.

9.2 Chemical Calculations

- Stoichiometric problems are solved using conversion factors derived from a balanced chemical equation.
- A conversion factor called a mole ratio relates the moles of a given substance to the moles of the desired substance.
- Units such as mass, volume of gases (at STP), and particles are converted to moles when working stoichiometry problems.

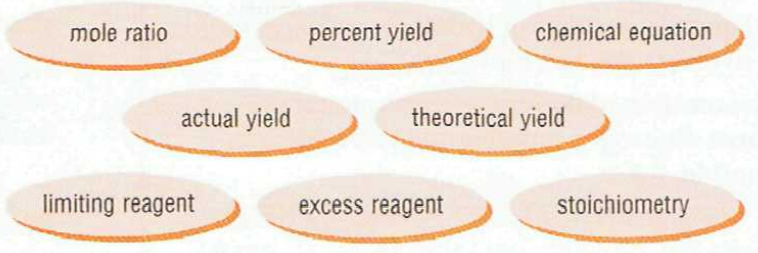
9.3 Limiting Reagent and Percent Yield

- Whenever quantities of two or more reactants are given in a stoichiometry problem, the limiting reagent must be identified.
- A limiting reagent is completely used up in a chemical reaction.
- The amount of limiting reagent determines the amount of product formed in a chemical reaction.
- If there is a single limiting reagent in a reaction, all the other reactants are in excess.
- A theoretical yield is the maximum amount of product that can be obtained from a given amount of reactants in a chemical reaction.
- An actual yield is the amount of product obtained when the reaction is carried out in the laboratory.
- A ratio of the actual yield to the theoretical yield, expressed as a percentage, is the percent yield of a reaction.

CHAPTER CONCEPT MAP

Use these terms to construct a concept map that organizes the major ideas of this chapter.

Chem ASAP! Concept Map 9
Create your Concept Map using the computer.



Chapter 9 REVIEW

CONCEPT PRACTICE

33. Interpret each chemical equation in terms of interacting particles. *9.1*

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

b. $4NH_3(g) + 6NO(g) \longrightarrow 5N_2(g) + 6H_2O(g)$

c. $4K(s) + O_2(g) \longrightarrow 2K_2O(s)$

34. Interpret each equation in Problem 33 in terms of interacting numbers of moles of reactants and products. *9.1*

35. Calculate and compare the mass of the reactants with the mass of the products for each equation in Problem 33. Show that each balanced equation obeys the law of conservation of mass. *9.1*

36. Explain the term mole ratio in your own words. When would you use this term? *9.2*

37. Carbon disulfide is an important industrial solvent. It is prepared by the reaction of coke with sulfur dioxide. *9.2*

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

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?

38. 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. *9.2*

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

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?

39. The reaction of fluorine with ammonia produces dinitrogen tetrafluoride and hydrogen fluoride. *9.2*

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

a. If you have 66.6 g NH_3, how many grams of F_2 are required for 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?

40. What information about a chemical reaction is derived from the coefficients in a balanced equation? *9.2*

41. Lithium nitride reacts with water to form ammonia and aqueous lithium hydroxide. *9.2*

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

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).

42. What is the significance of the limiting reagent in a chemical process? What happens to the amount of any reagent that is present in an excess? *9.3*

43. How would you identify a limiting reagent in a chemical reaction? *9.3*

44. For each balanced equation, identify the limiting reagent for the given combination of reactants. *9.3*

a. $2Al + 3Cl_2 \longrightarrow 2AlCl_3$ (3.6 mol Al, 5.3 mol Cl_2)

b. $2H_2 + O_2 \longrightarrow 2H_2O$ (6.4 mol H_2, 3.4 mol O_2)

c. $2P_2O_5 + 6H_2O \longrightarrow 4H_3PO_4$ (0.48 mol P_2O_5, 1.52 mol H_2O)

d. $4P + 5O_2 \longrightarrow 2P_2O_5$ (14.5 mol P, 18.0 mol O_2)

45. For each reaction in Problem 44, calculate the number of moles of product formed. *9.3*

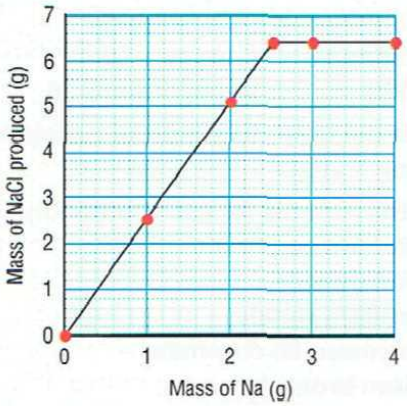
46. For each reaction in Problem 44, calculate the number of moles of excess reagent remaining after the reaction. *9.3*

47. 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? *9.3*

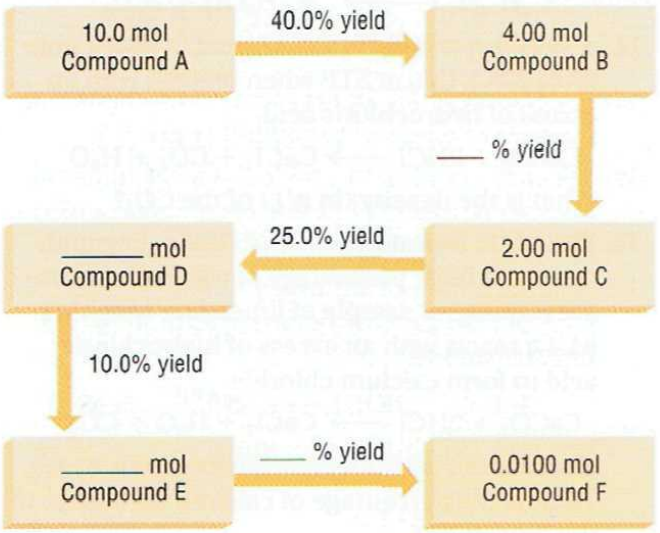
48. In an experiment, varying masses of sodium metal are reacted with a fixed initial mass of chlorine gas. The amounts of sodium used and the amounts of sodium chloride formed are shown on the following graph. *9.3*



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.

49. What does the percent yield of a chemical reaction measure? *9.3*

50. 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 forth. The diagram below summarizes the stepwise 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. *9.3*



CONCEPT MASTERY

51. 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.

52. 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.

53. 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 1.0 kg O_2?
b. How many grams of the excess reagent remain after the reaction?

54. 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?

55. 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?

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

56. 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.)

CRITICAL THINKING

57. Choose the term that completes the second relationship.

a. *equation : coefficients* — *balance:*_____
(1) moles (3) weight
(2) standard masses (4) atoms

b. actual : theoretical — experimental:_____
(1) excess (3) real
(2) limiting (4) calculated

58. 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%?

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

60. 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

CUMULATIVE REVIEW

61. 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.

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

63. *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. $H_2SO_4(aq) + Mg(OH)_2(aq) \longrightarrow MgSO_4(aq) + H_2O(l)$

64. Write a net ionic equation for each reaction in Problem 63.

65. Identify the spectator ions in each reaction in Problem 63.

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

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

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

68. Write the formula for each compound.

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

69. 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.

CONCEPT CHALLENGE

70. A car gets 9.2 kilometers to a liter of gasoline. *Assuming that gasoline is 100% octane* (C_8H_{18}), which has a specific gravity of 0.69, 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.

71. 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.)

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

72. 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?

73. The white limestone cliffs of Dover, England, contain a large percentage of calcium carbonate ($CaCO_3$). *A sample of limestone weighing* 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?

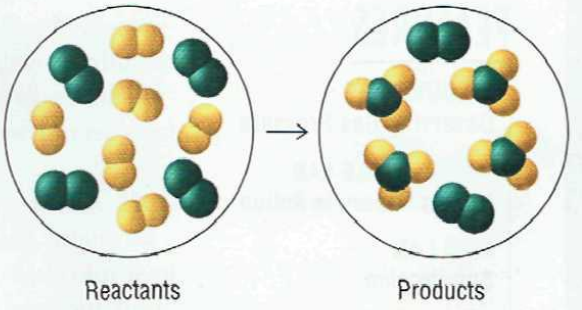
Chapter 9 STANDARDIZED TEST PREP

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 mol 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. Which of the following are needed to calculate the percent yield of a reaction?
   I. theoretical yield
   II. excess yield
   III. actual yield
   a. I only
   b. I and II only
   c. I and III only
   d. II and III only
   e. I, II, and III
4. According to the balanced equation for the combustion of ethane, C_2H_6,
   $2C_2H_6(g) + 7O_2(g) \longrightarrow 4CO_2(g) + 6H_2O(l)$,
   for every molecule of ethane that reacts,
   a. 2 molecules of carbon dioxide are produced.
   b. 3 molecules of water are produced.
   c. Both (a) and (b) are correct.
   d. Neither (a) nor (b) is correct.
5. 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 2.0 mol of excess N_2.
   b. 4.0 mol of Mg_3N_2 and 1.0 mol of excess Mg.
   c. 6.0 mol of Mg_3N_2 and 3.0 mol of excess N_2.
   d. no product because the reactants are not in the correct mole ratio.

Questions 6 and 7 involve the reaction between diatomic element P and diatomic element Q to form the compound P_3Q.

6. Write a balanced equation for the reaction between element P and element Q.
7. 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
8.	Every stoichiometry calculation uses a balanced equation.	BECAUSE	Every chemical reaction obeys the law of conservation of mass.
9.	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.
10.	The amount of the limiting reagent left after a reaction is zero.	BECAUSE	The limiting reagent is completely used up in a reaction.
11.	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.
12.	A mole ratio is always written with the larger number in the numerator.	BECAUSE	A mole ratio will always be greater than one.