3 Chemical Equations and Reaction Stoichiometry

OUTLINE

OBJECTIVES

After you have studied this chapter, you should be able to

- *Write balanced chemical equations to describe chemical reactions*
- *Interpret balanced chemical equations to calculate the* moles *of reactants and products involved in each of the reactions*
- *Interpret balanced chemical equations to calculate the* masses *of reactants and products involved in each of the reactions*
- *Determine which reactant is the limiting reactant in reactions*
- *Use the limiting reactant concept in calculations recording chemical equations*
- *Compare the amount of substance actually formed in a reaction (actual yield) with the predicted amount (theoretical yield), and determine the percent yield*
- *Work with sequential reactions*
- *Use the terminology of solutions—solute, solvent, concentration*
- *Calculate concentrations of solutions when they are diluted*
- *Carry out calculations related to the use of solutions in chemical reactions*

Steel wool burns in pure oxygen gas.

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

In Chapter 2 we studied composition stoichiometry, the quantitative relationships among elements in compounds. In this chapter as we study reaction stoichiometry—the quantitative relationships among substances as they participate in chemical reactions—we ask several important questions. *How* can we describe the reaction of one substance with another? *How much* of one substance reacts with a given amount of another substance? *Which reactant* determines the amounts of products formed in a chemical reaction? *How* can we describe reactions in aqueous solutions?

Whether we are concerned with describing a reaction used in a chemical analysis, one used industrially in the production of a plastic, or one that occurs during metabolism in the body, we must describe it accurately. Chemical equations represent a very precise, yet a very versatile, language that describes chemical changes. We begin our study by examining chemical equations.

Methane, CH_4, is the main component of natural gas.

3-1 CHEMICAL EQUATIONS

See the *Saunders Interactive General Chemistry CD-ROM,* Screens 4.1–4.4, Equations.

Chemical reactions always involve changing one or more substances into one or more different substances. In other words, chemical reactions rearrange atoms or ions to form other substances.

Chemical equations are used to describe chemical reactions, and they show (1) *the substances that react*, called **reactants;** (2) *the substances formed*, called **products;** and (3) *the relative amounts of the substances involved.* We write the reactants to the *left* of an arrow and the products to the *right* of the arrow. As a typical example, let's consider the combustion (burning) of natural gas, a reaction used to heat buildings and cook food. Natural gas is a mixture of several substances, but the principal component is methane, CH_4. The equation that describes the reaction of methane with excess oxygen is

Sometimes it is not possible to represent a chemical change with a single chemical equation. For example, when too little O_2 is present, both CO_2 and CO are found as products, and a second chemical equation must be used to describe the process. In the present case (excess oxygen), only one equation is required.

$$\underbrace{CH_4 + 2O_2}_{\text{reactants}} \longrightarrow \underbrace{CO_2 + 2H_2O}_{\text{products}}$$

What does this equation tell us? In the simplest terms, it tells us that methane reacts with oxygen to produce carbon dioxide, CO_2, and water. More specifically, it says that for every CH_4 molecule that reacts, two molecules of O_2 also react, and that one CO_2 molecule and two H_2O molecules are formed. That is,

The arrow may be read "yields." The capital Greek letter delta (Δ) is sometimes used in place of the word "heat."

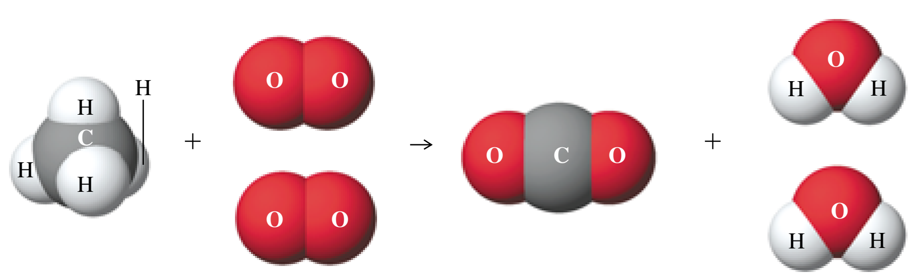
$$\underset{\text{1 molecule}}{CH_4} + \underset{\text{2 molecules}}{2O_2} \xrightarrow{\text{heat}} \underset{\text{1 molecule}}{CO_2} + \underset{\text{2 molecules}}{2H_2O}$$

This description of the reaction of CH_4 with O_2 is based on *experimental observations.* By this we mean experiments have shown that when one CH_4 molecule reacts with two O_2 molecules, one CO_2 molecule and two H_2O molecules are formed. *Chemical equations are based on experimental observations.* Special conditions required for some reactions are indicated by notation over the arrow. Figure 3-1 is a pictorial representation of the rearrangement of atoms described by this equation.

As we pointed out in Section 1-1, *there is no detectable change in the quantity of matter during an ordinary chemical reaction.* This guiding principle, the **Law of Conservation of Matter,** provides the basis for "balancing" chemical equations and for calculations based on those equations. Because matter is neither created nor destroyed during a chemical reaction,

> a balanced chemical equation must always include the same number of each kind of atom on both sides of the equation.

(a)



(b) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Figure 3-1 Two representations of the reaction of methane with oxygen to form carbon dioxide and water. Chemical bonds are broken and new ones are formed in each representation. Part (a) illustrates the reaction using models, and (b) uses chemical formulas.

Chemists usually write equations with the smallest possible whole-number coefficients.

Before we attempt to balance an equation, all substances must be represented by formulas that describe them *as they exist.* For instance, we must write H_2 to represent diatomic hydrogen molecules—not H, which represents hydrogen atoms. Once the correct formulas are written, the subscripts in the formulas may not be changed. Different subscripts in formulas specify different compounds, so changing the formulas would mean that the equation would no longer describe the same reaction.

Dimethyl ether, C_2H_6O, burns in an excess of oxygen to give carbon dioxide and water. Let's balance the equation for this reaction. In unbalanced form, the equation is

$$C_2H_6O + O_2 \longrightarrow CO_2 + H_2O$$

In Reactants	In Products
2C, 6H, 3O	1C, 2H, 3O

C, H are not balanced

Carbon appears in only one compound on each side, and the same is true for hydrogen. We begin by balancing these elements:

$$C_2H_6O + O_2 \longrightarrow \boxed{2CO_2} + \boxed{3H_2O}$$

In Reactants	In Products
2C, 6H, 3O	2C, 6H, 7O

Now O is not balanced

Now we have an odd number of atoms of O on each side. The single O in C_2H_6O balances one of the atoms of O on the right. We balance the other six by placing a coefficient of 3 before O_2 on the left.

$$C_2H_6O + \boxed{3O_2} \longrightarrow 2CO_2 + 3H_2O$$

In Reactants	In Products
2C, 6H, 7O	2C, 6H, 7O

Now the equation is balanced

When we think we have finished the balancing, we should *always* do a complete check for each element, as shown in red in the margin.

Let's generate the balanced equation for the reaction of aluminum metal with hydrochloric acid to produce aluminum chloride and hydrogen. The unbalanced "equation" is

$$Al + HCl \longrightarrow AlCl_3 + H_2$$

In Reactants	In Products
1Al, 1H, 1Cl	1Al, 2H, 3Cl

H, Cl are not balanced

As it now stands, the "equation" does not satisfy the Law of Conservation of Matter because there are two H atoms in the H_2 molecule and three Cl atoms in one formula unit of $AlCl_3$ (right side), but only one H atom and one Cl atom in the HCl molecule (left side).

Let us first balance chlorine by putting a coefficient of 3 in front of HCl.

$$Al + \boxed{3HCl} \longrightarrow AlCl_3 + H_2$$

In Reactants	In Products
1Al, 3H, 3Cl	1Al, 2H, 3Cl

Now H is not balanced

Now there are 3H on the left and 2H on the right. The least common multiple of 3 and 2 is 6; to balance H, we multiply the 3HCl by 2 and the H_2 by 3.

$$Al + \boxed{6HCl} \longrightarrow AlCl_3 + \boxed{3H_2}$$

In Reactants	In Products
1Al, 6H, 6Cl	1Al, 6H, 3Cl

Now Cl is not balanced

Now Cl is again unbalanced (6Cl on the left, 3 on the right), but we can fix this by putting a coefficient of 2 in front of $AlCl_3$ on the right.

$$Al + 6HCl \longrightarrow \boxed{2AlCl_3} + 3H_2$$

In Reactants	In Products
1Al, 6H, 6Cl	2Al, 6H, 6Cl

Now Al is not balanced

Now all elements except Al are balanced (1 on the left, 2 on the right); we complete the balancing by putting a coefficient of 2 in front of Al on the left.

$$\underset{\text{aluminum}}{\boxed{2Al}} + \underset{\text{hydrochloric acid}}{6HCl} \longrightarrow \underset{\text{aluminum chloride}}{2AlCl_3} + \underset{\text{hydrogen}}{3H_2}$$

In Reactants	In Products
2Al, 6H, 6Cl	2Al, 6H, 6Cl

Now the equation is balanced

Balancing chemical equations "by inspection" is a *trial-and-error* approach. It requires a great deal of practice, but it is *very important!* Remember that we use the smallest whole-number coefficients. Some chemical equations are difficult to balance by inspection or "trial and error." In Chapter 11 we will learn methods for balancing complex equations.

Problem-Solving Tip: *Balancing Chemical Equations*

There is no one best place to start when balancing a chemical equation, but the following suggestions might be helpful:

(1) Look for elements that appear in only one place on each side of the equation (in only one reactant and in only one product), and balance those elements first.

(2) If free, uncombined elements appear on either side, balance them last.

Notice how these suggestions worked in the procedures illustrated in this section. Above all, remember that we should *never* change subscripts in formulas, because doing so would describe different substances. We only adjust the coefficients to balance the equation.

3-2 CALCULATIONS BASED ON CHEMICAL EQUATIONS

See the *Saunders Interactive General Chemistry CD-ROM,* Screens 5.1–5.3, Stoichiometry.

We are now ready to use chemical equations to calculate the relative *amounts* of substances involved in chemical reactions. Let us again consider the combustion of methane in excess oxygen. The balanced chemical equation for that reaction is

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

On a quantitative basis, at the molecular level, the equation says

$$\underset{\substack{\text{1 molecule}\\ \text{of methane}}}{CH_4} + \underset{\substack{\text{2 molecules}\\ \text{of oxygen}}}{2O_2} \longrightarrow \underset{\substack{\text{1 molecule}\\ \text{of carbon dioxide}}}{CO_2} + \underset{\substack{\text{2 molecules}\\ \text{of water}}}{2H_2O}$$

EXAMPLE 3-1 *Number of Molecules*

How many O_2 molecules react with 47 CH_4 molecules according to the preceding equation?

Plan

A balanced chemical equation may be interpreted on a *molecular* basis.

The *balanced* equation tells us that *one* CH_4 molecule reacts with *two* O_2 molecules. We can construct two unit factors from this fact:

$$\frac{1\ CH_4\ \text{molecule}}{2\ O_2\ \text{molecules}} \quad \text{and} \quad \frac{2\ O_2\ \text{molecules}}{1\ CH_4\ \text{molecule}}$$

These expressions are unit factors for *this* reaction because the numerator and denominator are *chemically equivalent*. In other words, the numerator and the denominator represent the same amount of reaction. To convert CH_4 molecules to O_2 molecules, we multiply by the second of the two factors.

Solution

$$\underline{?}\ O_2\ \text{molecules} = 47\ CH_4\ \text{molecules} \times \frac{2\ O_2\ \text{molecules}}{1\ CH_4\ \text{molecule}} = \boxed{94\ O_2\ \text{molecules}}$$

You should now work Exercise 8.

We usually cannot work with individual molecules; a mole of a substance is an amount we might use in a laboratory experiment.

A chemical equation also indicates the relative amounts of each reactant and product in a given chemical reaction. We showed earlier that formulas can represent moles of substances. Suppose Avogadro's number of CH_4 molecules, rather than just one CH_4 molecule, undergo this reaction. Then the equation can be written

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 6.02\times 10^{23}\text{ molecules} & & 2(6.02\times 10^{23}\text{ molecules}) & & 6.02\times 10^{23}\text{ molecules} & & 2(6.02\times 10^{23}\text{molecules}) \\ =1\text{ mol} & & =2\text{ mol} & & =1\text{ mol} & & =2\text{ mol} \end{array}$$

This interpretation tells us that *one* mole of methane reacts with *two* moles of oxygen to produce *one* mole of carbon dioxide and *two* moles of water.

A balanced chemical equation may be interpreted in terms of *moles* of reactants and products.

Problem-Solving Tip: ***Are you Using a Balanced Equation?***

Never start a calculation involving a chemical reaction without first checking that the equation is balanced.

EXAMPLE 3-2 *Number of Moles Formed*

How many moles of water could be produced by the reaction of 3.5 mol of methane with excess oxygen (*i.e.*, more than a sufficient amount of oxygen is present)?

Plan

The equation for the combustion of methane

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \end{array}$$

shows that 1 mol of methane reacts with 2 mol of oxygen to produce 2 mol of water. From this information we construct two *unit factors:*

Please don't try to memorize unit factors for chemical reactions; rather, learn the general *method* for constructing them from balanced chemical equations.

$$\frac{1\text{ mol }CH_4}{2\text{ mol }H_2O} \quad \text{and} \quad \frac{2\text{ mol }H_2O}{1\text{ mol }CH_4}$$

We use the second factor in this calculation.

Solution

$$\underline{?}\text{ mol }H_2O = 3.5\text{ mol }CH_4 \times \frac{2\text{ mol }H_2O}{1\text{ mol }CH_4} = \boxed{7.0\text{ mol }H_2O}$$

You should now work Exercises 10 and 14.

We know the mass of 1 mol of each of these substances, so we can also write

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \\ 16.0\text{ g} & & 2(32.0\text{ g}) & & 44.0\text{ g} & & 2(18.0\text{ g}) \\ 16.0\text{ g} & & 64.0\text{ g} & & 44.0\text{ g} & & 36.0\text{ g} \\ \multicolumn{3}{c}{\underbrace{\hphantom{16.0\text{ g} \quad 64.0\text{ g}}}_{80.0\text{ g reactants}}} & & \multicolumn{3}{c}{\underbrace{\hphantom{44.0\text{ g} \quad 36.0\text{ g}}}_{80.0\text{ g products}}} \end{array}$$

The equation now tells us that 16.0 grams of CH_4 reacts with 64.0 grams of O_2 to form 44.0 grams of CO_2 and 36.0 grams of H_2O. The Law of Conservation of Matter is satisfied. Chemical equations describe **reaction ratios,** that is, the *mole ratios* of reactants and products as well as the *relative masses* of reactants and products.

A balanced equation may be interpreted on a *mass* basis.

Problem-Solving Tip: ***Use the Mole Ratio in Calculations with Balanced Chemical Equations***

The most important way of interpretating the balanced chemical equation is in terms of moles. We use the coefficients to get the mole ratio of any two substances we want to relate. Then we apply it as

$$\begin{matrix}\text{moles of}\\ \text{desired}\\ \text{substance}\end{matrix} = \begin{pmatrix}\text{moles of}\\ \text{substance}\\ \text{given}\end{pmatrix} \times \begin{pmatrix}\text{mole ratio}\\ \text{from balanced}\\ \text{chemical equation}\end{pmatrix}$$

It is important to include the substance formulas as part of the units; this can help us decide how to set up the unit factors. Notice that in Example 3-2, we want to cancel the term mol CH_4, so we know that mol CH_4 must be in the denominator of the mole ratio by which we multiply; we want mol O_2 as the answer, so mol O_2 must appear in the numerator of the mole ratio. In other words, do not just write $\frac{\text{mol}}{\text{mol}}$; write $\frac{\text{mol of something}}{\text{mol of something else}}$, giving the formulas of the two substances involved.

EXAMPLE 3-3 *Mass of a Reactant Required*

What mass of oxygen is required to react completely with 1.20 mol of CH_4?

Plan

The balanced equation

$$\begin{matrix} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \\ 16.0\text{ g} & & 2(32.0\text{ g}) & & 44.0\text{ g} & & 2(18.0\text{ g}) \end{matrix}$$

gives the relationships among moles and grams of reactants and products.

$$\text{mol } CH_4 \longrightarrow \text{mol } O_2 \longrightarrow \text{g } O_2$$

Solution

$$\underline{?}\text{ g } O_2 = 1.20\text{ mol } CH_4 \times \frac{2\text{ mol } O_2}{1\text{ mol } CH_4} \times \frac{32.0\text{ g } O_2}{1\text{ mol } O_2} = 76.8\text{ g } O_2$$

EXAMPLE 3-4 *Mass of a Reactant Required*

What mass of oxygen is required to react completely with 24.0 g of CH_4?

Plan

Recall the balanced equation in Example 3-3.

$$\begin{matrix} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \\ 16.0\text{ g} & & 2(32.0)\text{ g} & & 44.0\text{ g} & & 2(18.0)\text{ g} \end{matrix}$$

This shows that 1 mol of CH_4 reacts with 2 mol of O_2. These two quantities are chemically equivalent, so we can construct *unit factors.*

Solution

$$\underset{1\text{ mol}}{CH_4} + \underset{2\text{ mol}}{2O_2} \longrightarrow \underset{1\text{ mol}}{CO_2} + \underset{2\text{ mol}}{2H_2O}$$

Here we solve the problem in three steps and so we convert

$$\underline{?}\text{ mol } CH_4 = 24.0\text{ g } CH_4 \times \frac{1\text{ mol } CH_4}{16.0\text{ g } CH_4} = \underline{1.50\text{ mol } CH_4}$$

1. g CH_4 → mol CH_4

$$\underline{?}\text{ mol } O_2 = 1.50\text{ mol } CH_4 \times \frac{2\text{ mol } O_2}{1\text{ mol } CH_4} = \underline{3.00\text{ mol } O_2}$$

2. mol CH_4 → mol O_2

$$\underline{?}\text{ g } O_2 = 3.00\text{ mol } O_2 \times \frac{32.0\text{ g } O_2}{1\text{ mol } O_2} = \boxed{96.0\text{ g } O_2}$$

3. mol O_2 → g O_2

All these steps could be combined into one setup in which we convert

$$\boxed{\text{g of } CH_4} \longrightarrow \boxed{\text{mol of } CH_4} \longrightarrow \boxed{\text{mol of } O_2} \longrightarrow \boxed{\text{g of } O_2}$$

$$\underline{?}\text{ g } O_2 = 24.0\text{ g } CH_4 \times \frac{1\text{ mol } CH_4}{16.0\text{ g } CH_4} \times \frac{2\text{ mol } O_2}{1\text{ mol } CH_4} \times \frac{32.0\text{ g } O_2}{1\text{ mol } O_2} = \boxed{96.0\text{ g } O_2}$$

The same answer, 96.0 g of O_2, is obtained by both methods.

You should now work Exercise 18.

The question posed in Example 3-4 may be reversed, as in Example 3-5.

EXAMPLE 3-5 *Mass of a Reactant Required*

What mass of CH_4, in grams, is required to react with 96.0 grams of O_2?

Plan

We recall that one mole of CH_4 reacts with two moles of O_2.

Solution

$$\underline{?}\text{ g } CH_4 = 96.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{1\text{ mol } CH_4}{2\text{ mol } O_2} \times \frac{16.0\text{ g } CH_4}{1\text{ mol } CH_4} = \boxed{24.0\text{ g } CH_4}$$

These unit factors are the reciprocals of those used in Example 3-4.

You should now work Exercise 22.

This is the amount of CH_4 in Example 3-4 that reacted with 96.0 grams of O_2.

EXAMPLE 3-6 *Mass of a Product Formed*

Most combustion reactions occur in excess O_2, that is, more than enough O_2 to burn the substance completely. Calculate the mass of CO_2, in grams, that can be produced by burning 6.00 mol of CH_4 in excess O_2.

It is important to recognize that the reaction must stop when the 6.0 mol of CH_4 has been used up. Some O_2 will remain unreacted.

Plan

The balanced equation tells us that one mole of CH_4 produces one mole of CO_2.

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \\ 16.0\text{ g} & & 2(32.0\text{ g}) & & 44.0\text{ g} & & 2(18.0\text{ g}) \end{array}$$

Solution

$$\underline{?}\text{ g } CO_2 = 6.00\text{ mol } CH_4 \times \frac{1\text{ mol } CO_2}{1\text{ mol } CH_4} \times \frac{44.0\text{ g } CO_2}{1\text{ mol } CO_2} = \boxed{2.64 \times 10^2\text{ g } CO_2}$$

You should now work Exercise 20.

Reaction stoichiometry usually involves interpreting a balanced chemical equation to relate a *given* bit of information to the *desired* bit of information.

EXAMPLE 3-7 *Mass of a Reactant Required*

Phosphorus, P_4, burns with excess oxygen to form tetraphosphorus decoxide, P_4O_{10}. In this reaction, what mass of P_4 reacts with 1.50 moles of O_2?

Plan

The balanced equation tells us that one mole of P_4 reacts with five moles of O_2.

$$\begin{array}{ccccc} P_4 & + & 5O_2 & \longrightarrow & P_4O_{10} \\ 1\text{ mol} & & 5\text{ mol} & & 1\text{ mol} \end{array}$$

$$\boxed{\text{mol } O_2} \longrightarrow \boxed{\text{mol } P_4} \longrightarrow \boxed{\text{mass } P_4}$$

Solution

$$\underline{?}\text{ g } P_4 = 1.50\text{ mol } O_2 \times \frac{1\text{ mol } P_4}{5\text{ mol } O_2} \times \frac{124.0\text{ g } P_4}{\text{mol } P_4} = \boxed{37.2\text{ g } P_4}$$

You should now work Exercise 24.

The possibilities for this kind of problem solving are limitless. Before you continue, you should work Exercises 8–25 at the end of the chapter.

3-3 THE LIMITING REACTANT CONCEPT

See the *Saunders Interactive General Chemistry CD-ROM*, Screen 5.5, Limiting Reactants.

In the problems we have worked thus far, the presence of an excess of one reactant was stated or implied. The calculations were based on the substance that was used up first, called the **limiting reactant.** Before we study the concept of the limiting reactant in stoichiometry, let's develop the basic idea by considering a simple but analogous nonchemical example.

Suppose you have four slices of ham and six slices of bread and you wish to make as many ham sandwiches as possible using only one slice of ham and two slices of bread per sandwich. Obviously, you can make only three sandwiches, at which point you run out of bread. (In a chemical reaction this would correspond to one of the reactants being used up—so the reaction would stop.) The bread is therefore the "limiting reactant," and the extra slice of ham is the "excess reactant." The amount of product, ham sandwiches, is determined by the amount of the limiting reactant, bread in this case. The limiting reactant is not necessarily the reactant present in the smallest amount. We have four slices of ham, the smallest amount, and six slices of bread, but the "reaction ratio" is two slices of bread to one piece of ham, and so bread is the limiting reactant.

EXAMPLE 3-8 *Limiting Reactant*

What mass of CO_2 could be formed by the reaction of 16.0 g of CH_4 with 48.0 g of O_2?

Plan

The balanced equation tells us that one mole of CH_4 reacts with two moles of O_2.

$$\begin{array}{ccccccc} CH_4 & + & 2O_2 & \longrightarrow & CO_2 & + & 2H_2O \\ 1\text{ mol} & & 2\text{ mol} & & 1\text{ mol} & & 2\text{ mol} \\ 16.0\text{ g} & & 2(32.0\text{ g}) & & 44.0\text{ g} & & 2(18.0\text{ g}) \end{array}$$

We are given masses of both CH_4 and O_2, so we calculate the number of moles of each reactant, and then determine the number of moles of each reactant required to react with the other. From these calculations we identify the limiting reactant. Then we base the calculation on it.

Solution

$$\underline{?}\text{ mol } CH_4 = 16.0\text{ g } CH_4 \times \frac{1\text{ mol } CH_4}{16.0\text{ g } CH_4} = \underline{1.00\text{ mol } CH_4}$$

$$\underline{?}\text{ mol } O_2 = 48.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} = \underline{1.50\text{ mol } O_2}$$

Now we return to the balanced equation. First we calculate the number of moles of O_2 that would be required to react with 1.00 mole of CH_4.

$$\underline{?}\text{ mol } O_2 = 1.00\text{ mol } CH_4 \times \frac{2\text{ mol } O_2}{1\text{ mol } CH_4} = 2.00\text{ mol } O_2$$

We see that 2.00 moles of O_2 is required, but we have 1.50 moles of O_2, so O_2 is the limiting reactant. Alternatively, we can calculate the number of moles of CH_4 that would react with 1.50 moles of O_2.

$$\underline{?}\text{ mol } CH_4 = 1.50\text{ mol } O_2 \times \frac{1\text{ mol } CH_4}{2\text{ mol } O_2} = 0.750\text{ mol } CH_4$$

This tells us that only 0.750 mole of CH_4 would be required to react with 1.50 moles of O_2. But we have 1.00 mole of CH_4, so we see again that O_2 is the limiting reactant. The reaction must stop when the limiting reactant, O_2, is used up, so we base the calculation on O_2.

$$\text{g of } O_2 \longrightarrow \text{mol of } O_2 \longrightarrow \text{mol of } CO_2 \longrightarrow \text{g of } CO_2$$

$$\underline{?}\text{ g } CO_2 = 48.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\ O_2} \times \frac{1\text{ mol } CO_2}{2\text{ mol } O_2} \times \frac{44.0\text{ g } CO_2}{1\text{ mol } CO_2} = 33.0\text{ g } CO_2$$

You should now work Exercise 26.

Thus, 33.0 grams of CO_2 is the most CO_2 that can be produced from 16.0 grams of CH_4 and 48.0 grams of O_2. If we had based our calculation on CH_4 rather than O_2, our answer would be too big (44.0 grams) and *wrong* because more O_2 than we have would be required.

Another approach to problems like Example 3-8 is to calculate the number of moles of each reactant:

$$\underline{?} \text{ mol } CH_4 = 16.0 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4} = \underline{1.00 \text{ mol } CH_4}$$

$$\underline{?} \text{ mol } O_2 = 48.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.0 \text{ g } O_2} = \underline{1.50 \text{ mol } O_2}$$

Then we return to the balanced equation. We first calculate the *required ratio* of reactants as indicated by the balanced chemical equation. We then calculate the *available ratio* of reactants and compare the two:

Required Ratio	**Available Ratio**
$\frac{1 \text{ mol } CH_4}{2 \text{ mol } O_2} = \frac{0.500 \text{ mol } CH_4}{1.00 \text{ mol } O_2}$	$\frac{1.00 \text{ mol } CH_4}{1.50 \text{ mol } O_2} = \frac{0.667 \text{ mol } CH_4}{1.00 \text{ mol } O_2}$

We see that each mole of O_2 requires exactly 0.500 mol of CH_4 to be completely used up. We have 0.667 mol of CH_4 for each mole of O_2, so there is more than enough CH_4 to react with the O_2 present. That means that there is *insufficient* O_2 to react with all of the available CH_4. The reaction must stop when the O_2 is gone; O_2 is the limiting reactant, and we must base the calculation on it. (If the available ratio of CH_4 to O_2 had been *smaller than* the required ratio, we would have concluded that there is not enough CH_4 to react with all of the O_2, and CH_4 would have been the limiting reactant.)

✓ Problem-Solving Tip: *Choosing the Limiting Reactant*

Students often wonder how to know which ratio to calculate to help find the limiting reactant.

(1) The ratio must involve the two reactants with amounts given in the problem.

(2) It doesn't matter which way you calculate the ratio, as long as you calculate both required ratio and available ratio in the same order. For example, we could calculate the required and available ratio of $\frac{\text{mol } O_2}{\text{mol } CH_4}$ in the approach we just illustrated. If you can't decide how to solve a limiting reactant problem, as a last resort, do the entire calculation twice—once based on each reactant amount given. The *smaller* answer is the right one.

EXAMPLE 3-9 *Limiting Reactant*

What is the maximum mass of $Ni(OH)_2$ that could be prepared by mixing two solutions that contain 25.9 g of $NiCl_2$ and 10.0 g of NaOH, respectively?

$$NiCl_2 + 2NaOH \longrightarrow Ni(OH)_2 + 2NaCl$$

Plan

Interpreting the balanced equation as usual, we have

$$\begin{array}{cccc} NiCl_2 \; + & 2NaOH & \longrightarrow \; Ni(OH)_2 \; + & 2NaCl \\ \text{1 mol} & \text{2 mol} & \text{1 mol} & \text{2 mol} \\ \text{129.6 g} & \text{2(40.0 g)} & \text{92.7 g} & \text{2(58.4 g)} \end{array}$$

We determine the number of moles of $NiCl_2$ and NaOH present. Then we find the number of moles of each reactant required to react with the other reactant. These calculations identify the limiting reactant. We base the calculation on it.

Solution

$$\underline{?}\text{ mol NiCl}_2 = 25.9\text{ g NiCl}_2 \times \frac{1\text{ mol NiCl}_2}{129.6\text{ g NiCl}_2} = 0.200\text{ mol NiCl}_2$$

$$\underline{?}\text{ mol NaOH} = 10.0\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} = 0.250\text{ mol NaOH}$$

Even though the reaction occurs in aqueous solution, this calculation is similar to earlier examples because we are given the amounts of pure reactants.

We return to the balanced equation and calculate the number of moles of NaOH required to react with 0.200 mol of $NiCl_2$.

$$\underline{?}\text{ mol NaOH} = 0.200\text{ mol NiCl}_2 \times \frac{2\text{ mol NaOH}}{1\text{ mol NiCl}_2} = 0.400\text{ mol NaOH}$$

But we have only 0.250 mol of NaOH, so NaOH is the limiting reactant.

$$\text{g of NaOH} \longrightarrow \text{mol of NaOH} \longrightarrow \text{mol Ni(OH)}_2 \longrightarrow \text{g of Ni(OH)}_2$$

$$\underline{?}\text{ g Ni(OH)}_2 = 10.0\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} \times \frac{1\text{ mol Ni(OH)}_2}{2\text{ mol NaOH}} \times \frac{92.7\text{ g Ni(OH)}_2}{1\text{ mol Ni(OH)}_2}$$

$$= \boxed{11.6\text{ g Ni(OH)}_2}$$

You should now work Exercises 30 and 34.

A precipitate of solid $Ni(OH)_2$ forms when colorless NaOH solution is added to green $NiCl_2$ solution.

Problem-Solving Tip: ***How Can We Recognize a Limiting Reactant Problem?***

When we are given the amounts of *two* (*or more*) reactants, we should suspect that we are dealing with a limiting reactant problem. It is very unlikely that *exactly* the stoichiometric amounts of both reactants are present in a reaction mixture.

3-4 PERCENT YIELDS FROM CHEMICAL REACTIONS

The **theoretical yield** from a chemical reaction is the yield calculated by assuming that the reaction goes to completion. In practice we often do not obtain as much product from a reaction mixture as is theoretically possible. This is true for several reasons. (1) Many reactions do not go to completion; that is, the reactants are not completely converted to products. (2) In some cases, a particular set of reactants undergoes two or more reactions

In the examples we have worked to this point, the amounts of products that we calculated were theoretical yields.

simultaneously, forming undesired products as well as desired products. Reactions other than the desired one are called "side reactions." (3) In some cases, separation of the desired product from the reaction mixture is so difficult that not all of the product formed is successfully isolated. The **actual yield** is the amount of a specified pure product actually obtained from a given reaction.

The term **percent yield** is used to indicate how much of a desired product is obtained from a reaction.

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 5.6, Percent Yield.

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

Consider the preparation of nitrobenzene, $C_6H_5NO_2$, by the reaction of a limited amount of benzene, C_6H_6, with excess nitric acid, HNO_3. The balanced equation for the reaction may be written as

$$\begin{array}{ccccccc} C_6H_6 & + & HNO_3 & \longrightarrow & C_6H_5NO_2 & + & H_2O \\ \text{1 mol} & & \text{1 mol} & & \text{1 mol} & & \text{1 mol} \\ \text{78.1 g} & & \text{63.0 g} & & \text{123.1 g} & & \text{18.0 g} \end{array}$$

EXAMPLE 3-10 *Percent Yield*

A 15.6-gram sample of C_6H_6 is mixed with excess HNO_3. We isolate 18.0 grams of $C_6H_5NO_2$. What is the percent yield of $C_6H_5NO_2$ in this reaction?

Plan

First we interpret the balanced chemical equation to calculate the theoretical yield of $C_6H_5NO_2$. Then we use the actual (isolated) yield and the previous definition to calculate the percent yield.

Solution

We calculate the theoretical yield of $C_6H_5NO_2$.

It is not necessary to know the mass of one mole of HNO_3 to solve this problem.

$$\underline{?}\ \text{g } C_6H_5NO_2 = 15.6 \text{ g } C_6H_6 \times \frac{1 \text{ mol } C_6H_6}{78.1 \text{ g } C_6H_6} \times \frac{1 \text{ mol } C_6H_5NO_2}{1 \text{ mol } C_6H_6} \times \frac{123.1 \text{ g } C_6H_5NO_2}{1 \text{ mol } C_6H_5NO_2}$$

$$= 24.6 \text{ g } C_6H_5NO_2 \leftarrow \text{theoretical yield}$$

This tells us that if *all* the C_6H_6 were converted to $C_6H_5NO_2$ and isolated, we should obtain 24.6 grams of $C_6H_5NO_2$ (100% yield). We isolate only 18.0 grams of $C_6H_5NO_2$, however.

$$\text{percent yield} = \frac{\text{actual yield of product}}{\text{theoretical yield of product}} \times 100\% = \frac{18.0 \text{ g}}{24.6 \text{ g}} \times 100\%$$

$$= \boxed{73.2 \text{ percent yield}}$$

You should now work Exercise 38.

The amount of nitrobenzene obtained *in this experiment* is 73.2% of the amount that would be expected *if* the reaction had gone to completion, *if* there were no side reactions, and *if* we could have recovered all of the product as pure substance.

3-5 SEQUENTIAL REACTIONS

Often more than one reaction is required to change starting materials into the desired product. This is true for many reactions that we carry out in the laboratory and for many industrial processes. These are called **sequential reactions.** The amount of desired product from each reaction is taken as the starting material for the next reaction.

EXAMPLE 3-11 *Sequential Reactions*

At high temperatures, carbon reacts with water to produce a mixture of carbon monoxide, CO, and hydrogen, H_2.

$$C + H_2O \xrightarrow{\text{heat}} CO + H_2$$

Carbon monoxide is separated from H_2 and then used to separate nickel from cobalt by forming a gaseous compound, nickel tetracarbonyl, $Ni(CO)_4$.

$$Ni + 4CO \longrightarrow Ni(CO)_4$$

What mass of $Ni(CO)_4$ could be obtained from the CO produced by the reaction of 75.0 g of carbon? Assume 100% yield.

Plan

We interpret both chemical equations in the usual way, and solve the problem in two steps. They tell us that one mole of C produces one mole of CO and that four moles of CO is required to produce one mole of $Ni(CO)_4$.

1. We determine the number of moles of CO formed in the first reaction.
2. From the number of moles of CO produced in the first reaction, we calculate the number of grams of $Ni(CO)_4$ formed in the second reaction.

Solution

1.

$$\begin{array}{ccccccc} C & + & H_2O & \longrightarrow & CO & + & H_2 \\ \text{1 mol} & & \text{1 mol} & & \text{1 mol} & & \text{1 mol} \\ \text{12.0 g} & & & & & & \end{array}$$

$$\underline{?}\ \text{mol CO} = 75.0\ \text{g C} \times \frac{1\ \text{mol C}}{12.0\ \text{g C}} \times \frac{1\ \text{mol CO}}{1\ \text{mol C}} = 6.25\ \text{mol CO}$$

2.

$$\begin{array}{ccccc} Ni & + & 4CO & \longrightarrow & Ni(CO)_4 \\ \text{1 mol} & & \text{4 mol} & & \text{1 mol} \\ & & & & \text{171 g} \end{array}$$

$$\underline{?}\ \text{g Ni(CO)}_4 = 6.25\ \text{mol CO} \times \frac{1\ \text{mol Ni(CO)}_4}{4\ \text{mol CO}} \times \frac{171\ \text{g Ni(CO)}_4}{1\ \text{mol Ni(CO)}_4} = \boxed{267\ \text{g Ni(CO)}_4}$$

Alternatively, we can set up a series of unit factors based on the conversions in the reaction sequence and solve the problem in one setup.

$$\boxed{\text{g C}} \longrightarrow \boxed{\text{mol C}} \longrightarrow \boxed{\text{mol CO}} \longrightarrow \boxed{\text{mol Ni(CO)}_4} \longrightarrow \boxed{\text{g Ni(CO)}_4}$$

$$\underline{?}\ \text{g Ni(CO)}_4 = 75.0\ \text{g C} \times \frac{1\ \text{mol C}}{12.0\ \text{g C}} \times \frac{1\ \text{mol CO}}{1\ \text{mol C}} \times \frac{1\ \text{mol Ni(CO)}_4}{4\ \text{mol CO}} \times \frac{171\ \text{g Ni(CO)}_4}{1\ \text{mol Ni(CO)}_4} = \boxed{267\ \text{g Ni(CO)}_4}$$

You should now work Exercise 46.

EXAMPLE 3-12 *Sequential Reactions*

Two large uses for H_3PO_4 are in fertilizers and cola drinks. Approximately 100 lb of H_3PO_4– based fertilizer are used per year per person in America.

Phosphoric acid, H_3PO_4, is a very important compound used to make fertilizers. It is also present in cola drinks.

H_3PO_4 can be prepared in a two-step process.

$$\text{Rxn 1:}\quad P_4 + 5O_2 \longrightarrow P_4O_{10}$$

$$\text{Rxn 2:}\quad P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4$$

We allow 272 grams of phosphorus to react with excess oxygen, which forms tetraphosphorus decoxide, P_4O_{10}, in 89.5% yield. In the second step reaction, a 96.8% yield of H_3PO_4 is obtained. What mass of H_3PO_4 is obtained?

Plan

1. We interpret the first equation as usual and calculate the amount of P_4O_{10} *obtained.*

$$\begin{array}{ccccc} P_4 & + & 5O_2 & \longrightarrow & P_4O_{10} \\ 1\text{ mol} & & 5\text{ mol} & & 1\text{ mol} \\ 124\text{ g} & & 5(32.0\text{ g}) & & 284\text{ g} \end{array}$$

$$\boxed{\text{g } P_4} \longrightarrow \boxed{\text{mol } P_4} \longrightarrow \boxed{\text{mol } P_4O_{10}} \longrightarrow \boxed{\text{g } P_4O_{10}}$$

2. Then we interpret the second equation and calculate the amount of H_3PO_4 *obtained* from the P_4O_{10} from the first step.

$$\begin{array}{ccccc} P_4O_{10} & + & 6H_2O & \longrightarrow & 4H_3PO_4 \\ 1\text{ mol} & & 6\text{ mol} & & 4\text{ mol} \\ 284\text{ g} & & 6(18.0\text{ g}) & & 4(98.0\text{ g}) \end{array}$$

$$\boxed{\text{g } P_4O_{10}} \longrightarrow \boxed{\text{mol } P_4O_{10}} \longrightarrow \boxed{\text{mol } H_3PO_4} \longrightarrow \boxed{\text{g } H_3PO_4}$$

Solution

The unit factors that account for less than 100% reaction and less than 100% recovery are included in parentheses.

1. $$\underline{?}\text{ g } P_4O_{10} = 272\text{ g } P_4 \times \frac{1\text{ mol } P_4}{124\text{ g } P_4} \times \frac{1\text{ mol } P_4O_{10}}{1\text{ mol } P_4} \times \frac{284\text{ g } P_4O_{10}\text{ theoretical}}{1\text{ mol } P_4O_{10}\text{ theoretical}} \times \left(\frac{89.5\text{ g } P_4O_{10}\text{ actual}}{100\text{ g } P_4O_{10}\text{ theoretical}}\right) = \boxed{558\text{ g } P_4O_{10}}$$

2. $$\underline{?}\text{ g } H_3PO_4 = 558\text{ g } P_4O_{10} \times \frac{1\text{ mol } P_4O_{10}}{284\text{ g } P_4O_{10}} \times \frac{4\text{ mol } H_3PO_4}{1\text{ mol } P_4O_{10}} \times \frac{98.0\text{ g } H_3PO_4\text{ theoretical}}{1\text{ mol } H_3PO_4\text{ theoretical}} \times \left(\frac{96.8\text{ g } H_3PO_4\text{ actual}}{100\text{ g } H_3PO_4\text{ theoretical}}\right) = \boxed{746\text{ g } H_3PO_4}$$

You should now work Exercises 48 and 50.

Chemists have determined the structures of many naturally occurring compounds. One way of proving the structure of such a compound is by synthesizing it from available starting materials. Professor Grieco, now at Indiana University, was assisted by Majetich and Ohfune in the synthesis of helenalin, a powerful anticancer drug, in a 40-step process. This 40-step synthesis gave a remarkable average yield of about 90% for each step, which still resulted in an overall yield of only about 1.5%.

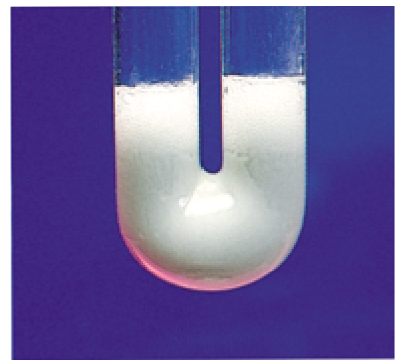
3-6 CONCENTRATIONS OF SOLUTIONS

Many chemical reactions are more conveniently carried out with the reactants mixed in solution rather than as pure substances. A **solution** is a homogeneous mixture, at the molecular level, of two or more substances. Simple solutions usually consist of one substance, the **solute,** dissolved in another substance, the **solvent.** The solutions used in the laboratory are usually liquids, and the solvent is often water. These are called **aqueous solutions.** For example, solutions of hydrochloric acid can be prepared by dissolving hydrogen chloride (HCl, a gas at room temperature and atmospheric pressure) in water. Solutions of sodium hydroxide are prepared by dissolving solid NaOH in water.

In some solutions, such as a nearly equal mixture of ethyl alcohol and water, the distinction between *solute* and *solvent* is arbitrary.

We often use solutions to supply the reactants for chemical reactions. Solutions allow the most intimate mixing of the reacting substances at the molecular level, much more than would be possible in solid form. (A practical example is drain cleaner, shown in the photo.) We sometimes adjust the concentrations of solutions to speed up or slow down the rate of a reaction. In this section we study methods for expressing the quantities of the various components present in a given amount of solution.

Concentrations of solutions are expressed in terms of *either* the amount of solute present in a given mass or volume of *solution*, or the amount of solute dissolved in a given mass or volume of *solvent.*



The sodium hydroxide and aluminum in some drain cleaners do not react while they are stored in solid form. When water is added, the NaOH dissolves and begins to act on trapped grease. At the same time, NaOH and Al react to produce H_2 gas; the resulting turbulence helps to dislodge the blockage. Do you see why the container should be kept tightly closed?

Percent by Mass

Concentrations of solutions may be expressed in terms of **percent by mass** of solute, which gives the mass of solute per 100 mass units of solution. The gram is the usual mass unit.

$$\text{percent solute} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

$$\text{percent} = \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100\%$$

Thus, a solution that is 10.0% calcium gluconate, $Ca(C_6H_{11}O_7)_2$, by mass contains 10.0 grams of calcium gluconate in 100.0 grams of *solution.* This could be described as 10.0 grams of calcium gluconate in 90.0 grams of water. The density of a 10.0% solution of calcium gluconate is 1.07 g/mL, so 100 mL of a 10.0% solution of calcium gluconate has a mass of 107 grams. Observe that 100 grams of a solution usually does *not* occupy 100 mL. Unless otherwise specified, percent means percent *by mass*, and water is the solvent.

A 10.0% solution of $Ca(C_6H_{11}O_7)_2$ is sometimes administered intravenously in emergency treatment for black widow spider bites.

EXAMPLE 3-13 *Percent of Solute*

Calculate the mass of nickel(II) sulfate, $NiSO_4$, contained in 200. g of a 6.00% solution of $NiSO_4$.

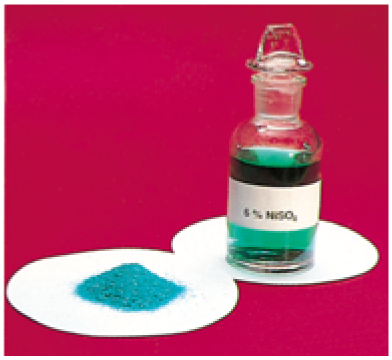
Plan

The percentage information tells us that the solution contains 6.00 grams of $NiSO_4$ per 100. grams of solution. The desired information is the mass of $NiSO_4$ in 200. grams of solution. A unit factor is constructed by placing 6.00 grams $NiSO_4$ over 100. grams of solution.

Multiplication of the mass of the solution, 200. grams, by this unit factor gives the mass of $NiSO_4$ in the solution.

Solution

$$\underline{?}\ \text{g } NiSO_4 = 200.\ \text{g soln} \times \frac{6.00\ \text{g } NiSO_4}{100.\ \text{g soln}} = \boxed{12.0\ \text{g } NiSO_4}$$



EXAMPLE 3-14 *Mass of Solution*

A 6.00% $NiSO_4$ solution contains 40.0 g of $NiSO_4$. Calculate the mass of the solution.

Plan

Placing 100. g of solution over 6.00 g of $NiSO_4$ gives the desired unit factor.

Solution

$$\underline{?}\ \text{g soln} = 40.0\ \text{g } NiSO_4 \times \frac{100.\ \text{g soln}}{6.00\ \text{g } NiSO_4} = \boxed{667\ \text{g soln}}$$

EXAMPLE 3-15 *Mass of Solute*

Calculate the mass of $NiSO_4$ present in 200. mL of a 6.00% solution of $NiSO_4$. The density of the solution is 1.06 g/mL at 25°C.

Plan

The volume of a solution multiplied by its density gives the mass of solution (see Section 1-11). The mass of solution is then multiplied by the mass fraction due to $NiSO_4$ (6.00 g $NiSO_4$/100. g soln) to give the mass of $NiSO_4$ in 200. mL of solution.

Solution

$$\underline{?}\ \text{g } NiSO_4 = \underbrace{200.\ \text{mL soln} \times \frac{1.06\ \text{g soln}}{1.00\ \text{mL soln}}}_{212\ \text{g soln}} \times \frac{6.00\ \text{g } NiSO_4}{100.\ \text{g soln}} = \boxed{12.7\ \text{g } NiSO_4}$$

You should now work Exercise 55.

EXAMPLE 3-16 *Percent Solute and Density*

What volume of a solution that is 15.0% iron(III) nitrate contains 30.0 g of $Fe(NO_3)_3$? The density of the solution is 1.16 g/mL at 25°C.

Plan

Two unit factors relate mass of $Fe(NO_3)_3$ and mass of solution, 15.0 g $Fe(NO_3)_3$/100 g and 100 g/15.0 g $Fe(NO_3)_3$. The second factor converts grams of $Fe(NO_3)_3$ to grams of solution.

Solution

$$\underline{?}\ \text{mL soln} = \underbrace{30.0\ \text{g } Fe(NO_3)_3 \times \frac{100.\ \text{g soln}}{15.0\ \text{g } Fe(NO_3)_3}}_{200\ \text{g soln}} \times \frac{1.00\ \text{mL soln}}{1.16\ \text{g soln}} = \boxed{172\ \text{mL}}$$

Note that the answer is not 200. mL but considerably less because 1.00 mL of solution has a mass of 1.16 grams; however, 172 mL of the solution has a mass of 200 g.

You should now work Exercise 58.

Molarity

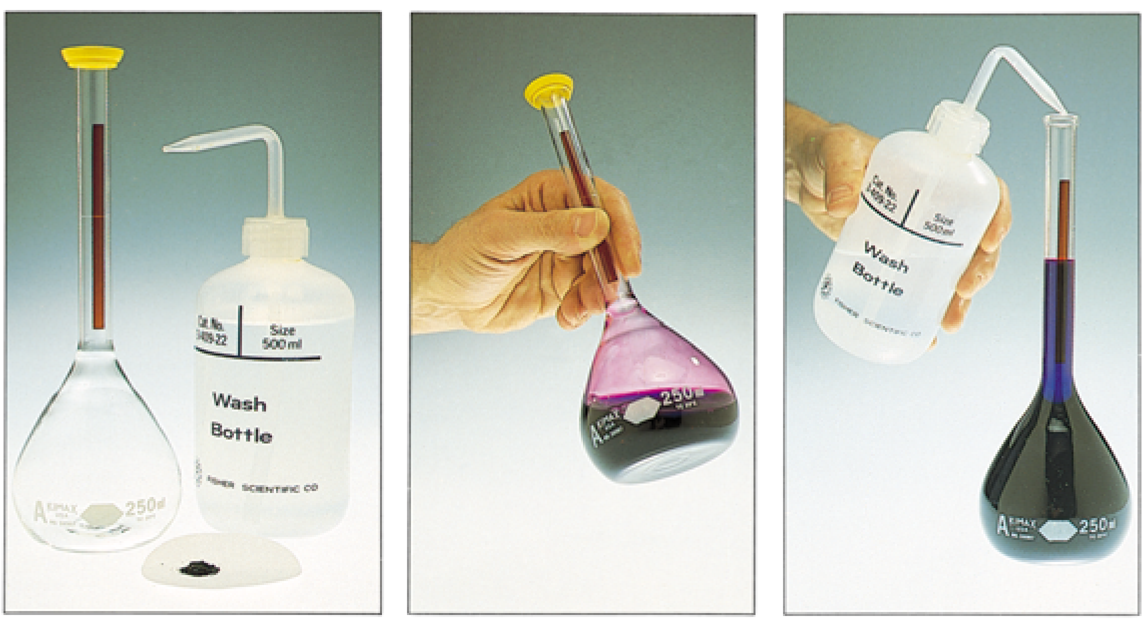
Molarity (M), or molar concentration, is a common unit for expressing the concentrations of solutions. **Molarity** is defined as the number of moles of solute per liter of solution:

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}}$$

To prepare one liter of a one molar solution, one mole of solute is placed in a one-liter volumetric flask, enough solvent is added to dissolve the solute, and solvent is then added until the volume of the solution is exactly one liter. Students sometimes make the mistake of assuming that a one molar solution contains one mole of solute in a liter of *solvent*. This is *not* the case; one liter of solvent *plus* one mole of solute usually has a total volume of more than one liter. A 0.100 M solution contains 0.100 mol of solute per liter of solution, and a 0.0100 M solution contains 0.0100 mol of solute per liter of solution (Figure 3-2).

See the *Saunders Interactive General Chemistry CD-ROM*, Screen 5.11, Preparing Solutions, Direct Addition.

The definition of molarity specifies the amount of solute *per unit volume of solution*, whereas percent specifies the amount of solute *per unit mass of solution*. Molarity therefore depends on temperature and pressure, whereas percent by mass does not.



(a) **(b)** **(c)**

Figure 3-2 Preparation of 0.0100 M solution of $KMnO_4$, potassium permanganate. A 250.-mL sample of 0.0100 M $KMnO_4$ solution contains 0.395 g of $KMnO_4$ (1 mol = 158 g). (a) 0.395 g of $KMnO_4$ (0.00250 mol) is weighed out carefully and transferred into a 250.-mL volumetric flask. (b) The $KMnO_4$ is dissolved in water. (c) Distilled H_2O is added to the volumetric flask until the volume of solution is 250. mL. The flask is then stoppered, and its contents are mixed thoroughly to give a homogeneous solution.

Water is the solvent in *most* of the solutions that we encounter. Unless otherwise indicated, we assume that water is the solvent. When the solvent is other than water, we state this explicitly.

EXAMPLE 3-17 *Molarity*

Calculate the molarity (M) of a solution that contains 3.65 grams of HCl in 2.00 liters of solution.

Plan

We are given the number of grams of HCl in 2.00 liters of solution. We apply the definition of molarity, remembering to convert grams of HCl to moles of HCl.

Solution

We place 3.65 g HCl over 2.00 L of solution, and then convert g HCl to mol HCl.

$$\frac{\underline{?}\text{ mol HCl}}{\text{L soln}} = \frac{3.65\text{ g HCl}}{2.00\text{ L soln}} \times \frac{1\text{ mol HCl}}{36.5\text{ g HCl}} = \boxed{0.0500\text{ mol HCl/L soln}}$$

The concentration of the HCl solution is 0.0500 molar, and the solution is called 0.0500 M hydrochloric acid. One liter of the solution contains 0.0500 mol of HCl.

You should now work Exercise 60.

EXAMPLE 3-18 *Mass of Solute*

Calculate the mass of $Ba(OH)_2$ required to prepare 2.50 L of a 0.0600 M solution of barium hydroxide.

Plan

The volume of the solution, 2.50 L, is multiplied by the concentration, 0.0600 mol $Ba(OH)_2$/L, to give the number of moles of $Ba(OH)_2$. The number of moles of $Ba(OH)_2$ is then multiplied by the mass of $Ba(OH)_2$ in one mole, 171.3 g $Ba(OH)_2$/mol $Ba(OH)_2$, to give the mass of $Ba(OH)_2$ in the solution.

Solution

$$\underline{?}\text{ g Ba(OH)}_2 = 2.50\text{ L soln} \times \frac{0.0600\text{ mol Ba(OH)}_2}{1\text{ L soln}} \times \frac{171.3\text{ g Ba(OH)}_2}{1\text{ mol Ba(OH)}_2}$$

$$= \boxed{25.7\text{ g Ba(OH)}_2}$$

You should now work Exercise 62.

The solutions of acids and bases that are sold commercially are too concentrated for most laboratory uses. We often dilute these solutions before we use them. We must know the molar concentration of a stock solution before it is diluted. This can be calculated from the specific gravity and the percentage data given on the label of the bottle.

EXAMPLE 3-19 *Molarity*

A sample of commercial sulfuric acid is 96.4% H_2SO_4 by mass, and its specific gravity is 1.84. Calculate the molarity of this sulfuric acid solution.

Plan

The density of a solution, grams per milliliter, is numerically equal to its specific gravity, so the density of the solution is 1.84 g/mL. The solution is 96.4% H_2SO_4 by mass; therefore 100. g of solution contains 96.4 g of *pure* H_2SO_4. From this information, we can find the molarity of the solution. First, we calculate the mass of one liter of solution.

Solution

$$\frac{\underline{?}\ \text{g soln}}{\text{L soln}} = \frac{1.84\ \text{g soln}}{\text{mL soln}} \times \frac{1000\ \text{mL soln}}{\text{L soln}} = 1.84 \times 10^3\ \text{g soln/L soln}$$

The solution is 96.4% H_2SO_4 by mass, so the mass of H_2SO_4 in one liter is

$$\frac{\underline{?}\ \text{g } H_2SO_4}{\text{L soln}} = \frac{1.84 \times 10^3\ \text{g soln}}{\text{L soln}} \times \frac{96.4\ \text{g } H_2SO_4}{100.0\ \text{g soln}} = 1.77 \times 10^3\ \text{g } H_2SO_4\text{/L soln}$$

The molarity is the number of moles of H_2SO_4 per liter of solution.

$$\frac{\underline{?}\ \text{mol } H_2SO_4}{\text{L soln}} = \frac{1.77 \times 10^3\ \text{g } H_2SO_4}{\text{L soln}} \times \frac{1\ \text{mol } H_2SO_4}{98.1\ \text{g } H_2SO_4} = \boxed{18.0\ \text{mol } H_2SO_4\text{/L soln}}$$

Thus, the solution is an 18.0 *M* H_2SO_4 solution. This problem can also be solved by using a series of three unit factors.

$$\frac{\underline{?}\ \text{mol } H_2SO_4}{\text{L soln}} = \frac{1.84\ \text{g soln}}{\text{mL soln}} \times \frac{1000\ \text{mL soln}}{\text{L soln}} \times \frac{96.4\ \text{g } H_2SO_4}{100\ \text{g soln}} \times \frac{1\ \text{mol } H_2SO_4}{98.1\ \text{g } H_2SO_4}$$

$$= 18.1\ \text{mol } H_2SO_4\text{/L soln} = \boxed{18.1\ M\ H_2SO_4}$$

You should now work Exercise 68.

ANALYSIS

Assay (H_2SO_4) W/W...Min. 95.0%--Max. 98.0%

MAXIMUM LIMITS OF IMPURITIES

Appearance........Passes A.C.S. Test
Color (APHA)....................10 Max.
Residue after Ignition...........4 ppm
Chloride (Cl)....................0.2 ppm
Nitrate (NO_3)..................0.5 ppm
Ammonium (NH_4)................1 ppm
Substances Reducing $KMnO_4$ (limit about 2ppm as SO_2).........Passes A.C.S. Test
Arsenic (As)..................0.004 ppm
Heavy Metals (as Pb)..........0.8 ppm
Iron (Fe)........................0.2 ppm
Mercury (Hg)......................5 ppb
Specific Gravity..................~1.84
Normality............................~36

Suitable for Mercury Determinations

A label that shows the analysis of sulfuric acid.

The small difference is due to rounding.

Problem-Solving Tip: *Write Complete Units*

A common pitfall is to write units that are not complete enough to be helpful. For instance, writing the density in Example 3-19 as just $\frac{1.84\ \text{g}}{\text{mL}}$ doesn't help us figure out the required conversions. It is much safer to write $\frac{1.84\ \text{g soln}}{\text{mL soln}}$, $\frac{1000\ \text{mL soln}}{\text{L soln}}$, $\frac{96.4\ \text{g } H_2SO_4}{100\ \text{g soln}}$, and so on. In Example 3-19, we have written complete units to help guide us through the problem.

3-7 DILUTION OF SOLUTIONS

Recall that the definition of molarity is the number of moles of solute divided by the volume of the solution in liters:

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}}$$

Multiplying both sides of the equation by the volume, we obtain

$$\text{volume (in L)} \times \text{molarity} = \text{number of moles of solute}$$

See the *Saunders Interactive General Chemistry CD-ROM,* Screen 5.12, Preparing Solutions, Dilution.

Multiplication of the volume of a solution, in liters, by its molar concentration gives the amount of solute in the solution.

A can of frozen orange juice contains a certain mass (or moles) of vitamin C. After the frozen contents of the can are diluted by addition of water, the amount of vitamin C in the resulting total amount of solution will be unchanged. The concentration, or amount per a selected volume, will be less in the final solution, however.

When we dilute a solution by mixing it with more solvent, the amount of solute present does not change. But the volume and the concentration of the solution *do* change. Because the same number of moles of solute is divided by a larger number of liters of solution, the molarity decreases. Using a subscript 1 to represent the original concentrated solution and a subscript 2 to represent the dilute solution, we obtain

$$\text{volume}_1 \times \text{molarity}_1 = \text{number of moles of solute} = \text{volume}_2 \times \text{molarity}_2$$

or

We could use any volume unit as long as we use the same unit on both sides of the equation. This relationship also applies when the concentration is changed by evaporating some solvent.

$$V_1M_1 = V_2M_2 \qquad \text{(for dilution only)}$$

This expression can be used to calculate any one of four quantities when the other three are known (Figure 3-3). We frequently need a certain volume of dilute solution of a given molarity for use in the laboratory, and we know the concentration of the initial solution available. Then we can calculate the amount of initial solution that must be used to make the dilute solution.

CAUTION! Dilution of a concentrated solution, especially of a strong acid or base, frequently liberates a great deal of heat. This can vaporize drops of water as they hit the concentrated solution and can cause dangerous spattering. As a safety precaution, *concentrated solutions of acids or bases are always poured slowly into water,* allowing the heat to be absorbed by the larger quantity of water. Calculations are usually simpler to visualize, however, by assuming that the water is added to the concentrated solution.

Example 3-20 *Dilution*

How many milliliters of 18.0 M H_2SO_4 are required to prepare 1.00 L of a 0.900 M solution of H_2SO_4?

Plan

The volume (1.00 L) and molarity (0.900 M) of the final solution, as well as the molarity (18.0 M) of the original solution, are given. Therefore, the relationship $V_1M_1 = V_2M_2$ can be used, with subscript 1 for the initial acid solution and subscript 2 for the dilute solution. We solve

$$V_1M_1 = V_2M_2 \qquad \text{for } V_1$$

Solution

$$V_1 = \frac{V_2M_2}{M_1} = \frac{1.00\text{ L} \times 0.900\ M}{18.0\ M} = 0.0500\text{ L} = \boxed{50.0\text{ mL}}$$

The dilute solution contains 1.00 L × 0.900 M = 0.900 mol of H_2SO_4, so 0.900 mol of H_2SO_4 must also be present in the original concentrated solution. Indeed, 0.0500 L × 18.0 M = 0.900 mol of H_2SO_4.

You should now work Exercises 70 and 72.

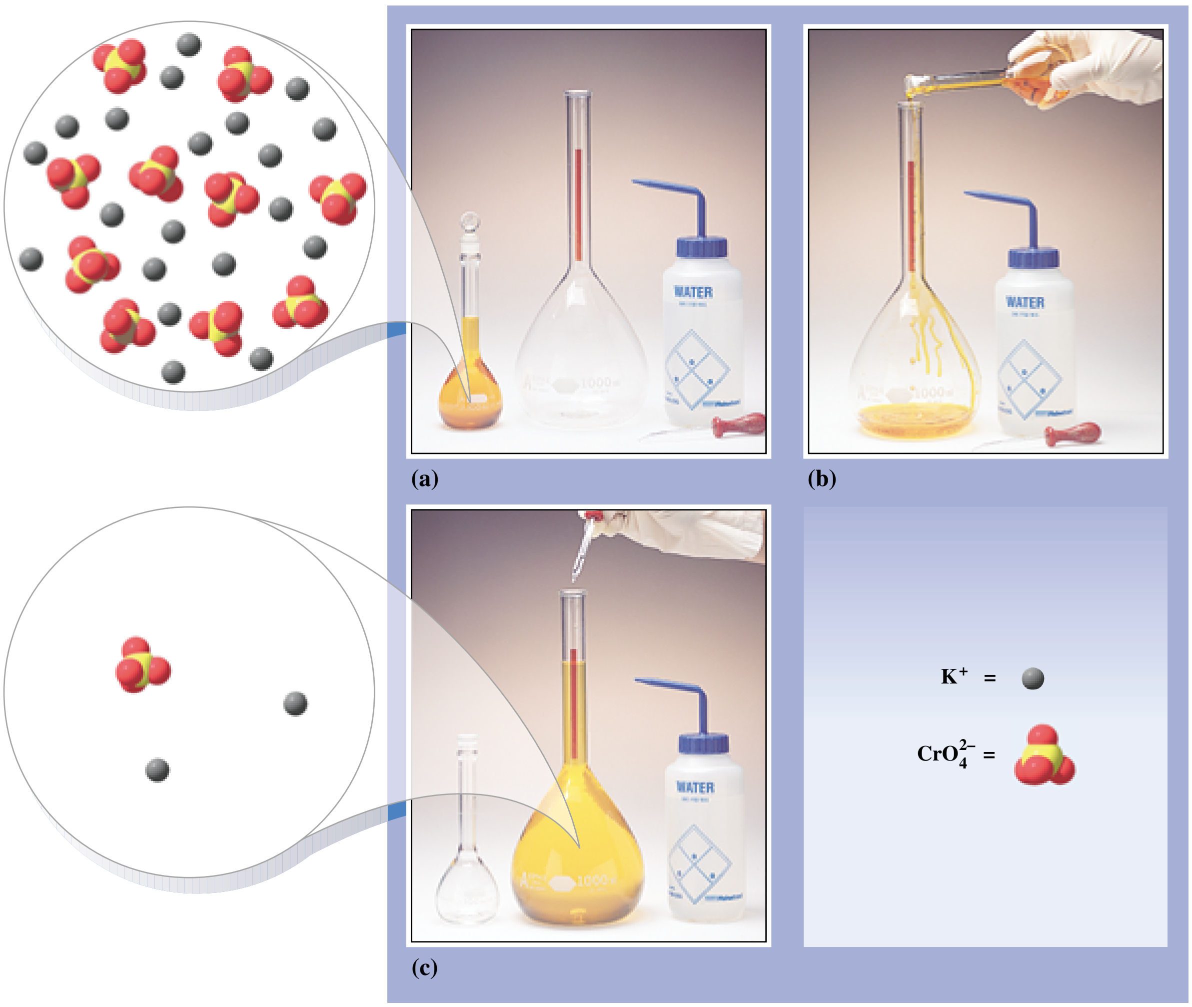


Figure 3-3 Dilution of a solution. (a) A 100.-mL volumetric flask is filled to the calibration line with a 0.100 M potassium chromate, K_2CrO_4 solution. (b) The 0.100 M K_2CrO_4 solution is transferred into a 1.00-L volumetric flask. The smaller flask is rinsed with a small amount of distilled H_2O. The rinse solution is added to the solution in the larger flask. To make sure that all the original K_2CrO_4 solution is transferred to the larger flask, the smaller flask is rinsed twice more and each rinse is added to the solution in the larger flask. (c) Distilled water is added to the 1.00-L flask until the liquid level coincides with its calibration line. The flask is stoppered and its contents are mixed thoroughly. The new solution is 0.0100 M K_2CrO_4. (100. mL of 0.100 M K_2CrO_4 solution has been diluted to 1000. mL.) The 100. mL of original solution and the 1000. mL of final solution both contain the amount of K_2CrO_4 dissolved in the original 100. mL of 0.100 M K_2CrO_4.

3-8 USING SOLUTIONS IN CHEMICAL REACTIONS

See the *Saunders Interactive General Chemistry CD-ROM,* Screens 5.13 and 5.15, Stoichiometry of Reactions in Solution.

If we plan to carry out a reaction in a solution, we must calculate the amounts of solutions that we need. If we know the molarity of a solution, we can calculate the amount of solute contained in a specified volume of that solution. This procedure is illustrated in Example 3-21.

EXAMPLE 3-21 *Amount of Solute*

Calculate (a) the number of moles of H_2SO_4 and (b) the number of grams of H_2SO_4 in 500. mL of 0.324 M H_2SO_4 solution.

Plan

Because we have two parallel calculations in this example, we state the plan for each step just before the calculation is done.

Solution

500. mL is more conveniently expressed as 0.500 L in this problem. By now, you should be able to convert mL to L (and the reverse) without writing out the conversion.

(a) The volume of a solution in liters multiplied by its molarity gives the number of moles of solute, H_2SO_4 in this case.

$$\underline{?}\text{ mol } H_2SO_4 = 0.500\text{ L soln} \times \frac{0.324\text{ mol } H_2SO_4}{\text{L soln}} = \boxed{0.162\text{ mol } H_2SO_4}$$

(b) We may use the results of part (a) to calculate the mass of H_2SO_4 in the solution.

A mole of H_2SO_4 is 98.1 g.

$$\underline{?}\text{ g } H_2SO_4 = 0.162\text{ mol } H_2SO_4 \times \frac{98.1\text{ g } H_2SO_4}{1\text{ mol } H_2SO_4} = \boxed{15.9\text{ g } H_2SO_4}$$

The mass of H_2SO_4 in the solution can be calculated without solving explicitly for the number of moles of H_2SO_4.

$$\underline{?}\text{ g } H_2SO_4 = 0.500\text{ L soln} \times \frac{0.324\text{ mol } H_2SO_4}{\text{L soln}} \times \frac{98.1\text{ g } H_2SO_4}{1\text{ mol } H_2SO_4} = \boxed{15.9\text{ g } H_2SO_4}$$

One of the most important uses of molarity relates the volume of a solution of known concentration of one reactant to the mass of the other reactant.

EXAMPLE 3-22 *Solution Stoichiometry*

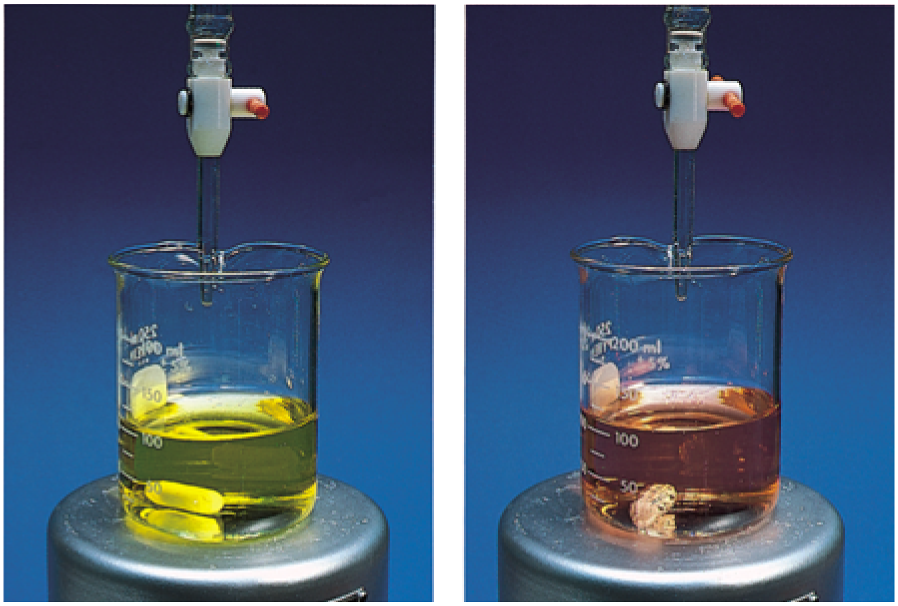
Calculate the volume in liters and in milliliters of a 0.324 M solution of sulfuric acid required to react completely with 2.792 grams of Na_2CO_3 according to the equation

$$H_2SO_4 + Na_2CO_3 \longrightarrow Na_2SO_4 + CO_2 + H_2O$$

Plan

The balanced equation tells us that one mole of H_2SO_4 reacts with one mole of Na_2CO_3, and we can write

$$\begin{array}{ccccccccc} H_2SO_4 & + & Na_2CO_3 & \longrightarrow & Na_2SO_4 & + & CO_2 & + & H_2O \\ 1\text{ mol} & & 1\text{ mol} & & 1\text{ mol} & & 1\text{ mol} & & 1\text{ mol} \\ & & 106.0\text{ g} & & & & & & \end{array}$$



The indicator methyl orange changes from yellow, its color in basic solutions, to orange, its color in acidic solutions, when the reaction in Example 3-22 reaches completion.

We convert (1) grams of Na_2CO_3 to moles of Na_2CO_3, (2) moles of Na_2CO_3 to moles of H_2SO_4, and (3) moles of H_2SO_4 to liters of H_2SO_4 solution.

$$\text{g } Na_2CO_3 \longrightarrow \text{mol } Na_2CO_3 \longrightarrow \text{mol } H_2SO_4 \longrightarrow \text{L } H_2SO_4 \text{ soln}$$

Solution

$$\underline{?}\ \text{L } H_2SO_4 = 2.792 \text{ g } Na_2CO_3 \times \frac{1 \text{ mol } Na_2CO_3}{106.0 \text{ g } Na_2CO_3} \times \frac{1 \text{ mol } H_2SO_4}{1 \text{ mol } Na_2CO_3} \times \frac{1 \text{ L } H_2SO_4 \text{ soln}}{0.324 \text{ mol } H_2SO_4}$$

$$= 0.0813 \text{ L } H_2SO_4 \text{ soln} \quad \text{or} \quad 81.3 \text{ mL } H_2SO_4 \text{ soln}$$

You should now work Exercise 76.

Often we must calculate the volume of solution of known molarity that is required to react with a specified volume of another solution. We always examine the balanced chemical equation for the reaction to determine the *reaction ratio*, that is, the relative numbers of moles of reactants.

Example 3-23 *Volume of Solution Required*

Find the volume in liters and in milliliters of a 0.505 M NaOH solution required to react with 40.0 mL of 0.505 M H_2SO_4 solution according to the reaction

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

Plan

We shall work this example in several steps, stating the "plan," or reasoning, just before each step in the calculation. Then we shall use a single setup to solve the problem.

Solution

The balanced equation tells us that the reaction ratio is 1 mol of H_2SO_4 to 2 mol of NaOH.

$$\underset{1\text{ mol}}{H_2SO_4} + \underset{2\text{ mol}}{2NaOH} \longrightarrow \underset{1\text{ mol}}{Na_2SO_4} + \underset{2\text{ mol}}{2H_2O}$$

From the volume and the molarity of the H_2SO_4 solution, we can calculate the number of moles of H_2SO_4.

The volume of H_2SO_4 solution is expressed as 0.0400 L rather than 40.0 mL.

$$\underline{?}\text{ mol } H_2SO_4 = 0.0400\text{ L } H_2SO_4\text{ soln} \times \frac{0.505\text{ mol } H_2SO_4}{\text{L soln}} = 0.0202\text{ mol } H_2SO_4$$

The number of moles of H_2SO_4 is related to the number of moles of NaOH by the reaction ratio, 1 mol H_2SO_4/2 mol NaOH:

$$\underline{?}\text{ mol NaOH} = 0.0202\text{ mol } H_2SO_4 \times \frac{2\text{ mol NaOH}}{1\text{ mol } H_2SO_4} = 0.0404\text{ mol NaOH}$$

Now we can calculate the volume of 0.505 *M* NaOH solution that contains 0.0404 mol of NaOH:

Again we see that molarity is a unit factor. In this case,

$$\frac{1.00\text{ L NaOH soln}}{0.505\text{ mol NaOH}}$$

$$\underline{?}\text{ L NaOH soln} = 0.0404\text{ mol NaOH} \times \frac{1.00\text{ L NaOH soln}}{0.505\text{ mol NaOH}} = \boxed{0.0800\text{ L NaOH soln}}$$

which we usually call 80.0 mL of NaOH solution.

We have worked through the problem stepwise; let us solve it in a single setup.

$$\boxed{\begin{array}{c}\text{L } H_2SO_4\text{ soln}\\ \text{available}\end{array}} \longrightarrow \boxed{\begin{array}{c}\text{mol } H_2SO_4\\ \text{available}\end{array}} \longrightarrow \boxed{\begin{array}{c}\text{mol NaOH}\\ \text{soln needed}\end{array}} \longrightarrow \boxed{\begin{array}{c}\text{L NaOH}\\ \text{soln needed}\end{array}}$$

$$\underline{?}\text{ L NaOH soln} = 0.0400\text{ L } H_2SO_4\text{ soln} \times \frac{0.505\text{ mol } H_2SO_4}{\text{L } H_2SO_4\text{ soln}} \times \frac{2\text{ mol NaOH}}{1\text{ mol } H_2SO_4} \times \frac{1.00\text{ L NaOH soln}}{0.505\text{ mol NaOH}}$$

$$= \boxed{0.0800\text{ L NaOH soln or } 80.0\text{ mL NaOH soln}}$$

You should now work Exercise 78.

Key Terms

Actual yield The amount of a specified pure product actually obtained from a given reaction. Compare with *Theoretical yield.*

Chemical equation Description of a chemical reaction by placing the formulas of reactants on the left and the formulas of products on the right of an arrow. A chemical equation must be balanced; that is, it must have the same number of each kind of atom on both sides.

Concentration The amount of solute per unit volume or mass of solvent or of solution.

Dilution The process of reducing the concentration of a solute in solution, usually simply by adding more solvent.

Limiting reactant A substance that stoichiometrically limits the amount of product(s) that can be formed.

Molarity (*M*) The number of moles of solute per liter of solution.

Percent by mass 100% multiplied by the mass of a solute divided by the mass of the solution in which it is contained.

Percent yield 100% times actual yield divided by theoretical yield.

Products Substances produced in a chemical reaction.

Reactants Substances consumed in a chemical reaction.

Reaction ratio The relative amounts of reactants and products involved in a reaction; may be the ratio of moles, or masses.

Reaction stoichiometry Description of the quantitative relationships among substances as they participate in chemical reactions.

Sequential reaction A chemical process in which several reaction steps are required to convert starting materials into products.

Solute The dispersed (dissolved) phase of a solution.

Solution A homogeneous mixture of two or more substances.

Solvent The dispersing medium of a solution.

Stoichiometry Description of the quantitative relationships among elements and compounds as they undergo chemical changes.

Theoretical yield The maximum amount of a specified product that could be obtained from specified amounts of reactants, assuming complete consumption of the limiting reactant according to only one reaction and complete recovery of the product. Compare with *Actual yield.*

Exercises

Chemical Equations

1. What is a chemical equation? What information does it contain?

2. When balancing chemical equations, you make certain that the same number of atoms of each element are on both sides of the equation. What scientific (natural) law requires that there be equal numbers of atoms of each element in both the products and the reactants?

3. Use words to state explicitly the relationships among numbers of molecules of reactants and products in the equation for the combustion of hexane, C_6H_{14}.

$$2C_6H_{14} + 19O_2 \longrightarrow 12CO_2 + 14H_2O$$

Balance each "equation" in Exercises 4–7 by inspection.

4. (a) $Al + Cl_2 \longrightarrow Al_2Cl_6$
(b) $N_2 + H_2 \longrightarrow NH_3$
(c) $K + KNO_3 \longrightarrow K_2O + N_2$
(d) $H_2O + KO_2 \longrightarrow KOH + O_2$
(e) $H_2SO_4 + NH_3 \longrightarrow (NH_4)_2SO_4$

5. (a) $P_4 + O_2 \longrightarrow P_4O_6$
(b) $P_4 + O_2 \longrightarrow P_4O_{10}$
(c) $K_2CO_3 + Al_2Cl_6 \longrightarrow Al_2(CO_3)_3 + KCl$
(d) $KClO_3 + C_{12}H_{22}O_{11} \longrightarrow KCl + CO_2 + H_2O$
(e) $KOH + H_3PO_4 \longrightarrow KH_2PO_4 + H_2O$

6. (a) $Fe_2O_3 + CO \longrightarrow Fe + CO_2$
(b) $Mg_3N_2 + H_2O \longrightarrow NH_3 + Mg(OH)_2$
(c) $Ca_3(PO_4)_2 + H_2SO_4 \longrightarrow Ca(H_2PO_4)_2 + Ca(HSO_4)_2$
(d) $(NH_4)_2Cr_2O_7 \longrightarrow N_2 + H_2O + Cr_2O_3$
(e) $Al + Cr_2O_3 \longrightarrow Al_2O_3 + Cr$

7. (a) $UO_2 + HF \longrightarrow UF_4 + H_2O$
(b) $NaCl + H_2O + SiO_2 \longrightarrow HCl + Na_2SiO_3$
(c) $Ca(HCO_3)_2 + Na_2CO_3 \longrightarrow CaCO_3 + NaHCO_3$
(d) $NH_3 + O_2 \longrightarrow NO + H_2O$
(e) $PCl_3 + O_2 \longrightarrow POCl_3$

Calculations Based on Chemical Equations

In Exercises 8–11, (a) write the balanced chemical equation that represents the reaction described by words, and then perform calculations to answer parts (b) and (c).

8. (a) Nitrogen, N_2, combines with hydrogen, H_2, to form ammonia, NH_3.
(b) How many hydrogen molecules are required to react with 600 nitrogen molecules?
(c) How many ammonia molecules are formed in part (b)?

9. (a) Sulfur, S_8, combines with oxygen at elevated temperatures to form sulfur dioxide.
(b) If 250 oxygen molecules are used up in this reaction, how many sulfur molecules react?
(c) How many sulfur dioxide molecules are formed in part (b)?

10. (a) Lime, CaO, dissolves in muriatic acid, HCl, to form calcium chloride, $CaCl_2$, and water.
(b) How many moles of HCl are required to dissolve 6.7 mol of CaO?
(c) How many moles of water are formed in part (b)?

11. (a) Aluminum building materials have a hard, transparent, protective coating of aluminum oxide, Al_2O_3, formed by reaction with oxygen in the air. The sulfuric acid, H_2SO_4, in acid rain dissolves this protective coating and forms aluminum sulfate, $Al_2(SO_4)_3$, and water.
(b) How many moles of H_2SO_4 are required to react with 6.8 mol of Al_2O_3?
(c) How many moles of $Al_2(SO_4)_3$ are formed in part (b)?

12. Calculate the number of grams of baking soda, $NaHCO_3$, that contain 14.0 moles of carbon.

13. Limestone, coral, and seashells are composed primarily of calcium carbonate. The test for the identification of a carbonate is to use a few drops of hydrochloric acid. The unbalanced equation is

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

(a) Balance the equation.
(b) How many atoms are in 0.250 moles of calcium carbonate?
(c) What number of carbon dioxide molecules is released on the reaction of 0.250 moles of calcium carbonate?

14. How many moles of oxygen can be obtained by the decomposition of 8.0 mol of reactant in each of the following reactions?
(a) $2KClO_3 \longrightarrow 2KCl + 3O_2$
(b) $2H_2O_2 \longrightarrow 2H_2O + O_2$
(c) $2HgO \longrightarrow 2Hg + O_2$
(d) $2NaNO_3 \longrightarrow 2NaNO_2 + O_2$
(e) $KClO_4 \longrightarrow KCl + 2O_2$

15. For the formation of 8.0 mol of water, which reaction uses the most nitric acid?
(a) $3Cu + 8HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$
(b) $Al_2O_3 + 6HNO_3 \longrightarrow 2Al(NO_3)_3 + 3H_2O$
(c) $4Zn + 10HNO_3 \longrightarrow 4Zn(NO_3)_2 + NH_4NO_3 + 3H_2O$

16. Consider the reaction

$$NH_3 + O_2 \xrightarrow{\text{not balanced}} NO + H_2O$$

For every 25.00 mol of NH_3, (a) how many moles of O_2 are required, (b) how many moles of NO are produced, and (c) how many moles of H_2O are produced?

17. What masses of cobalt(II) chloride and of hydrogen fluoride are needed to prepare 15.0 mol of cobalt(II) fluoride by the following reaction?

$$CoCl_2 + 2HF \longrightarrow CoF_2 + 2HCl$$

18. We allow 24.0 g of methane, CH_4, to react as completely as possible with excess oxygen, O_2, to form CO_2 and water. Write the balanced equation for this reaction. What mass of oxygen reacts?

***19.** Calculate the mass of calcium required to react with 3.770 g of carbon during the production of calcium carbide, CaC_2.

20. Sodium iodide, NaI, is a source of iodine used to produce iodized salt.
(a) Write the balanced chemical equation for the reaction of sodium and iodine.
(b) How many grams of sodium iodide are produced by the reaction of 93.25 grams of iodine?

21. Consider the reaction

$$2NO + Br_2 \longrightarrow 2NOBr$$

For every 7.50 mol of bromine that reacts, how many moles of (a) NO react and (b) NOBr are produced?

22. A sample of magnetic iron oxide, Fe_3O_4, reacts completely with hydrogen at red heat. The water vapor formed by the reaction

$$Fe_3O_4 + 4H_2 \xrightarrow{\text{heat}} 3Fe + 4H_2O$$

is condensed and found to weigh 18.75 g. Calculate the mass of Fe_3O_4 that reacted.

23. Iron(III) oxide, Fe_2O_3, is a result of the reaction of iron with the oxygen in air.
(a) What is the balanced equation for this reaction?
(b) What number of moles of iron react with 15.25 mol of oxygen from the air?
(c) What mass of iron is required to react with 15.25 mol of oxygen?

24. Calculate the number of molecules of propane, C_3H_8, that will produce 4.80 grams of water when burned in excess oxygen, O_2.

25. What mass of pentane, C_5H_{12}, produces 4.52×10^{22} CO_2 molecules when burned in excess oxygen, O_2?

Limiting Reactant

26. How many grams of NH_3 can be prepared from 59.85 g of N_2 and 12.11 g of H_2?

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

***27.** Silver nitrate solution reacts with calcium chloride solution according to the equation

$$2AgNO_3 + CaCl_2 \longrightarrow Ca(NO_3)_2 + 2AgCl$$

All of the substances involved in this reaction are soluble in water except silver chloride, AgCl, which forms a solid (precipitate) at the bottom of the flask. Suppose we mix together a solution containing 12.6 g of $AgNO_3$ and one containing 8.40 g of $CaCl_2$. What mass of AgCl is formed?

***28.** "Superphosphate," a water-soluble fertilizer, is sometimes marketed as "triple phosphate." It is a mixture of $Ca(H_2PO_4)_2$ and $CaSO_4$ on a 1:2 *mole* basis. It is formed by the reaction

$$Ca_3(PO_4)_2 + 2H_2SO_4 \longrightarrow Ca(H_2PO_4)_2 + 2CaSO_4$$

We treat 300 g of $Ca_3(PO_4)_2$ with 200 g of H_2SO_4. How many grams of superphosphate could be formed?

29. Gasoline is produced from crude oil, a nonrenewable resource. Ethanol is mixed with gasoline to produce a fuel called gasohol. Calculate the mass of water produced when

66.89 g of ethanol, C_2H_5OH, is burned in 55.21 g of oxygen.

30. What mass of potassium can be produced by the reaction of 125.0 g of Na with 125.0 g of KCl?

$$Na + KCl \xrightarrow{heat} NaCl + K$$

31. Silicon carbide, an abrasive, is made by the reaction of silicon dioxide with graphite.

$$SiO_2 + C \xrightarrow{heat} SiC + CO \qquad \text{(balanced?)}$$

We mix 452 g of SiO_2 and 306 g of C. If the reaction proceeds as far as possible, which reactant is left over? How much of this reactant remains?

32. Octane, C_8H_{18}, is a component of gasoline. A spark is used to ignite 1.563 grams of octane and 6.778 grams of oxygen in a sealed container. What mass of carbon dioxide is produced?

33. What mass of $Ca(NO_3)_2$ can be prepared by the reaction of 18.9 g of HNO_3 with 7.4 g of $Ca(OH)_2$?

$$2HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2H_2O$$

34. What is the maximum amount of $Ca_3(PO_4)_2$ that can be prepared from 7.4 g of $Ca(OH)_2$ and 9.8 g of H_3PO_4?

$$3Ca(OH)_2 + 2H_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6\ H_2O$$

35. A reaction mixture contains 11.0 g of PCl_3 and 7.00 g of PbF_2. What mass of $PbCl_2$ can be obtained from the following reaction?

$$3PbF_2 + 2PCl_3 \longrightarrow 2PF_3 + 3PbCl_2$$

How much of which reactant is left unchanged?

Percent Yield from Chemical Reactions

36. The percent yield for the reaction

$$PCl_3 + Cl_2 \longrightarrow PCl_5$$

is 83.2%. What mass of PCl_5 is expected from the reaction of 73.7 g of PCl_3 with excess chlorine?

37. The percent yield for the following reaction carried out in carbon tetrachloride solution is 59.0%.

$$Br_2 + Cl_2 \longrightarrow 2BrCl$$

(a) What amount of BrCl is formed from the reaction of 0.0250 mol Br_2 with 0.0250 mol Cl_2?
(b) What amount of Br_2 is left unchanged?

38. Solid silver nitrate undergoes thermal decomposition to form silver metal, nitrogen dioxide, and oxygen. Write the chemical equation for this reaction. A 0.443-g sample of silver metal is obtained from the decomposition of a 0.784-g sample of $AgNO_3$. What is the percent yield of the reaction?

39. Tin(IV) chloride is produced in 85.0% yield by the reaction of tin with chlorine. How much tin is required to produce a kilogram of tin(IV) chloride?

40. Ethylene oxide, C_2H_4O, a fumigant sometimes used by exterminators, is synthesized in 88.1% yield by reaction of ethylene bromohydrin, C_2H_5OBr, with sodium hydroxide:

$$C_2H_5OBr + NaOH \longrightarrow C_2H_4O + NaBr + H_2O$$

How many grams of ethylene bromohydrin are consumed in the production of 353 g of ethylene oxide, at 88.1% yield?

41. If 15.0 g sodium carbonate is obtained from the thermal decomposition of 75.0 g of sodium hydrogen carbonate,

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

what is the percent yield?

42. What is the percent yield if 112 mg SO_2 is obtained from the combustion of 78.1 mg of carbon disulfide according to the reaction

$$CS_2 + 3O_2 \longrightarrow CO_2 + 2SO_2?$$

43. From a 60.0-g sample of an iron ore containing Fe_3O_4, 2.09 g of Fe is obtained by the reaction

$$Fe_3O_4 + 2C \longrightarrow 3Fe + 2CO_2$$

What is the percent of Fe_3O_4 in the ore?

44. The reaction of finely divided aluminum and iron(III) oxide, Fe_2O_3, is called the thermite reaction. It produces a tremendous amount of heat, making the welding of railroad track possible. The reaction of 750. grams of aluminum and 750. grams of iron(III) oxide produces 247.5 grams of iron.
(a) Calculate the mass of iron that should be released by this reaction.
(b) What is the percent yield of iron?

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

45. Lime, $Ca(OH)_2$, can be used to neutralize an acid spill. A 5.06-g sample of $Ca(OH)_2$ reacts with an excess of hydrochloric acid; 6.74 g of calcium chloride is collected. What is the percent yield of this experiment?

$$Ca(OH)_2 + 2HCl \longrightarrow CaCl_2 + 2H_2O$$

Sequential Reactions

46. Consider the two-step process for the formation of tellurous acid described by the following equations:

$$TeO_2 + 2OH^- \longrightarrow TeO_3^{2-} + H_2O$$
$$TeO_3^{2-} + 2H^+ \longrightarrow H_2TeO_3$$

What mass of H_2TeO_3 is formed from 95.2 g of TeO_2, assuming 100% yield?

47. Consider the formation of cyanogen, C_2N_2, and its subsequent decomposition in water given by the equations

$$2Cu^{2+} + 6CN^- \longrightarrow 2[Cu(CN)_2]^- + C_2N_2$$
$$C_2N_2 + H_2O \longrightarrow HCN + HOCN$$

How much hydrocyanic acid, HCN, can be produced from 35.00 g of KCN, assuming 100% yield?

48. What mass of potassium chlorate is required to supply the proper amount of oxygen needed to burn 44.2 g of methane, CH_4?

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$
$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

49. Hydrogen, obtained by the electrical decomposition of water, is combined with chlorine to produce 84.2 g of hydrogen chloride. Calculate the mass of water decomposed.

$$2H_2O \longrightarrow 2H_2 + O_2$$
$$H_2 + Cl_2 \longrightarrow 2HCl$$

50. Ammonium nitrate, known for its use in agriculture, can be produced from ammonia by the following sequence of reactions:

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$$
$$NO(g) + O_2(g) \longrightarrow NO_2(g)$$
$$NO_2(g) + H_2O(\ell) \longrightarrow HNO_3(aq) + NO(g)$$
$$HNO_3(aq) + NH_3(g) \longrightarrow NH_4NO_3(aq)$$

(a) Balance each equation.
(b) How many moles of nitrogen atoms are required for every mole of ammonium nitrate (NH_4NO_3)?
(c) How much ammonia is needed to prepare 100.0 grams of ammonium nitrate (NH_4NO_3)?

51. Calcium sulfate is the essential component of plaster and sheet rock. Waste calcium sulfate can be converted into quicklime, CaO, by reaction with carbon at high temperatures. The following two reactions represent a sequence of reactions that might take place:

$$CaSO_4(s) + 4C(s) \longrightarrow CaS(\ell) + 4CO(g)$$
$$CaS(\ell) + 3CaSO_4(s) \longrightarrow 4CaO(s) + 4SO_2(g)$$

What weight of sulfur dioxide (in grams) could be obtained from 1.000 kg of calcium sulfate?

***52.** The Grignard reaction is a two-step reaction used to prepare pure hydrocarbons. Consider the preparation of pure ethane, CH_3CH_3, from ethyl chloride, CH_3CH_2Cl.

Step 1: $CH_3CH_2Cl + Mg \longrightarrow CH_3CH_2MgCl$

Step 2: $CH_3CH_2MgCl + H_2O \longrightarrow CH_3CH_3 + Mg(OH)Cl$

We allow 27.2 g of CH_3CH_2Cl (64.4 g/mol) to react with excess magnesium. From the first step reaction, CH_3CH_2MgCl (88.7 g/mol) is obtained in 79.5% yield. In the second step reaction, a 85.7% yield of CH_3CH_3 (30.0 g/mol) is obtained. What mass of CH_3CH_3 is obtained?

***53.** When sulfuric acid dissolves in water, the following reactions take place:

$$H_2SO_4 \longrightarrow H^+ + HSO_4^-$$
$$HSO_4^- \longrightarrow H^+ + SO_4^{2-}$$

The first reaction is 100.0% complete, and the second reaction is 10.0% complete. Calculate the concentrations of the various ions in a 0.150 M aqueous solution of H_2SO_4.

***54.** The chief ore of zinc is the sulfide, ZnS. The ore is concentrated by flotation and then heated in air, which converts the ZnS to ZnO.

$$2ZnS + 3O_2 \longrightarrow 2ZnO + 2SO_2$$

The ZnO is then treated with dilute H_2SO_4

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

to produce an aqueous solution containing the zinc as $ZnSO_4$. An electric current is passed through the solution to produce the metal.

$$2ZnSO_4 + 2H_2O \longrightarrow 2Zn + 2H_2SO_4 + O_2$$

What mass of Zn is obtained from an ore containing 225 kg of ZnS? Assume the flotation process to be 90.6% efficient, the electrolysis step to be 92.2% efficient, and the other steps to be 100% efficient.

Concentrations of Solutions—Percent by Mass

55. (a) How many moles of solute are contained in 500 g of a 2.00% aqueous solution of $K_2Cr_2O_7$?
(b) How many grams of solute are contained in the solution of part (a)?
(c) How many grams of water (the solvent) are contained in the solution of part (a)?

56. The density of an 18.0% solution of ammonium sulfate, $(NH_4)_2SO_4$, is 1.10 g/mL. What mass of $(NH_4)_2SO_4$ is required to prepare 275 mL of this solution?

57. The density of an 18.0% solution of ammonium chloride, NH_4Cl, solution is 1.05 g/mL. What mass of NH_4Cl does 275 mL of this solution contain?

58. What volume of the solution of $(NH_4)_2SO_4$ described in Exercise 56 contains 90.0 g of $(NH_4)_2SO_4$?

***59.** A reaction requires 33.6 g of NH_4Cl. What volume of the solution described in Exercise 57 do you need if you wish to use a 25.0% excess of NH_4Cl?

Concentrations of Solutions—Molarity

60. What is the molarity of a solution that contains 555 g of phosphoric acid, H_3PO_4, in 3.00 L of solution?

61. What is the molarity of a solution that contains 4.50 g of sodium chloride in 40.0 mL of solution?

62. How many grams of the cleansing agent Na_3PO_4 (a) are needed to prepare 250 mL of 0.50 *M* solution, and (b) are in 250 mL of 0.50 *M* solution?

63. How many kilograms of ethylene glycol, $C_2H_6O_2$, are needed to prepare a 9.50 *M* solution to protect a 14.0-L car radiator against freezing? What is the mass of $C_2H_6O_2$ in 14.0 L of 9.50 *M* solution?

64. A solution made by dissolving 16.0 g of $CaCl_2$ in 64.0 g of water has a density of 1.180 g/mL at 20°C.
(a) What is the percent by mass of $CaCl_2$ in the solution?
(b) What is the molarity of $CaCl_2$ in the solution?

65. A solution contains 0.100 mol/L of each of the following acids: HCl, H_2SO_4, H_3PO_4.
(a) Is the molarity the same for each acid?
(b) Is the number of molecules per liter the same for each acid?
(c) Is the mass per liter the same for each acid?

66. What is the molarity of a barium chloride solution prepared by dissolving 1.50 g of $BaCl_2 \cdot 2H_2O$ in enough water to make 600 mL of solution?

67. How many grams of potassium benzoate trihydrate, $KC_7H_5O_2 \cdot 3H_2O$, are needed to prepare 1 L of a 0.250 *M* solution of potassium benzoate?

68. Stock hydrofluoric acid solution is 49.0% HF and has a specific gravity of 1.17. What is the molarity of the solution?

69. Stock phosphoric acid solution is 85.0% H_3PO_4 and has a specific gravity of 1.70. What is the molarity of the solution?

Dilution of Solutions

70. Commercial concentrated hydrochloric acid is 12.0 *M* HCl. What volume of concentrated hydrochloric acid is required to prepare 4.50 L of 2.25 *M* HCl solution?

71. Commercially available concentrated sulfuric acid is 18.0 *M* H_2SO_4. Calculate the volume of concentrated sulfuric acid required to prepare 4.50 L of 2.25 *M* H_2SO_4 solution.

72. Calculate the volume of 0.0600 *M* $Ba(OH)_2$ solution that contains the same number of moles of $Ba(OH)_2$ as 225 mL of 0.0900 *M* $Ba(OH)_2$ solution.

73. Calculate the volume of 4.00 *M* NaOH solution required to prepare 200 mL of a 0.800 *M* solution of NaOH.

74. Calculate the volume of pure water that dilutes 100. mL of 12 *M* NaOH to 4.75 *M*.

75. In the laboratory preparation room is a reagent bottle that contains 5.0 L of 12 *M* NaOH. Write a set of instructions for the production of 250 mL of 3.0 *M* NaOH from the 12 *M* solution.

Using Solutions in Chemical Reactions

76. Calculate the volume of a 0.225 *M* solution of potassium hydroxide, KOH, required to react with 0.215 g of acetic acid, CH_3COOH, according to the following reaction.

$$KOH + CH_3COOH \longrightarrow KCH_3COO + H_2O$$

77. Calculate the number of grams of carbon dioxide, CO_2, that can react with 135 mL of a 0.357 *M* solution of potassium hydroxide, KOH, according to the following reaction.

$$2KOH + CO_2 \longrightarrow K_2CO_3 + H_2O$$

78. What volume of 0.246 *M* HNO_3 solution is required to react completely with 38.6 mL of 0.0515 *M* $Ba(OH)_2$?

$$Ba(OH)_2 + 2HNO_3 \longrightarrow Ba(NO_3)_2 + 2H_2O$$

79. What volume of 0.55 *M* HBr is required to react completely with 0.80 mol of $Ca(OH)_2$?

$$2HBr + Ca(OH)_2 \longrightarrow CaBr_2 + 2H_2O$$

80. An excess of $AgNO_3$ reacts with 185.5 mL of an $AlCl_3$ solution to give 0.325 g of AgCl. What is the concentration, in moles per liter, of the $AlCl_3$ solution?

$$AlCl_3 + 3AgNO_3 \longrightarrow 3AgCl + Al(NO_3)_3$$

81. An impure sample of solid Na_2CO_3 is allowed to react with 0.1755 *M* HCl.

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$$

A 0.2337-g sample of sodium carbonate requires 15.55 mL of HCl solution. What is the purity of the sodium carbonate?

82. Calculate the volume of 12 *M* HCl that just reacts with 15 g of aluminum.

83. What volume of 18 *M* H_2SO_4 do you need to measure in order to neutralize 2 L of 5.35 *M* NaOH? The products are sodium sulfate and water.

84. What volume of 0.0496 *M* $HClO_4$ reacts with 35.9 mL of a 0.505 *M* KOH solution according to the following reaction?

$$KOH + HClO_4 \longrightarrow KClO_4 + H_2O$$

***85.** What is the molarity of a solution of sodium hydroxide, NaOH, if 36.9 mL of this solution is required to react with 29.2 mL of 0.101 *M* hydrochloric acid solution according to the following reaction?

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

***86.** What is the molarity of a solution of sodium hydroxide, NaOH, if 29.8 mL of this solution is required to react with 25.0 mL of 0.0513 *M* nitric acid solution according to the following reaction?

$$HNO_3 + NaOH \longrightarrow NaNO_3 + H_2O$$

Mixed Exercises

87. What mass of sulfuric acid can be obtained from 1.00 kg of sulfur by the following series of reactions?

$$S_8 + 8O_2 \xrightarrow{98\%\ \text{yield}} 8SO_2$$

$$2SO_2 + O_2 \xrightarrow{96\%\ \text{yield}} 2SO_3$$

$$SO_3 + H_2SO_4 \xrightarrow{100\%\ \text{yield}} H_2S_2O_7$$

$$H_2S_2O_7 + H_2O \xrightarrow{97\%\ \text{yield}} 2H_2SO_4$$

***88.** What is the total mass of products formed when 38.8 g of carbon disulfide is burned in air? What mass of carbon disulfide must be burned to produce a mixture of carbon dioxide and sulfur dioxide that has a mass of 54.2 g?

$$CS_2 + 3O_2 \xrightarrow{\text{heat}} CO_2 + 2SO_2$$

***89.** Iron(II) chloride, $FeCl_2$, reacts with ammonia, NH_3, and water, H_2O, to produce iron(II) hydroxide, $Fe(OH)_2$, and ammonium chloride, NH_4Cl.
(a) Write the balanced equation for this reaction.
(b) We mix 78.5 g $FeCl_2$, 25.0 g NH_3, and 25.0 g H_2O, which then react as completely as possible. Which is the limiting reactant?
(c) How many grams of ammonium chloride, NH_4Cl, are formed?
(d) How many grams of each of the two leftover reactants remain at the completion of the reaction?

***90.** An iron ore that contains Fe_3O_4 reacts according to the reaction

$$Fe_3O_4 + 2C \longrightarrow 3Fe + 2CO_2$$

We obtain 2.09 g of Fe from the reaction of 55.0 g of the ore. What is the percent Fe_3O_4 in the ore?

***91.** If 86.3% of the iron can be recovered from an ore that is 43.2% magnetic iron oxide, Fe_3O_4, what mass of iron could be recovered from 2.00 kg of this ore? The reduction of magnetic iron oxide is a complex process that can be represented in simplified form as

$$Fe_3O_4 + 4CO \longrightarrow 3Fe + 4CO_2$$

92. Gaseous chlorine will displace bromide ion from an aqueous solution of potassium bromide to form aqueous potassium chloride and aqueous bromine. Write the chemical equation for this reaction. What mass of bromine is produced if 0.381 g of chlorine undergoes reaction?

93. Calculate the volume of 2.50 *M* phosphoric acid solution necessary to react with 45.0 mL of 0.150 *M* $Mg(OH)_2$.

$$2H_3PO_4 + 3Mg(OH)_2 \longrightarrow Mg_3(PO_4)_2 + 6H_2O$$

CONCEPTUAL EXERCISES

94. Using your own words, give a definition of a chemical reaction.

95. Magnesium burns with a bright, white flame in air.
(a) Write and balance the equation for the combination of magnesium and oxygen.
(b) Explain the meaning of the numbers used to balance the equation.
(c) How is a balanced equation an expression of the Law of Conservation of Matter?

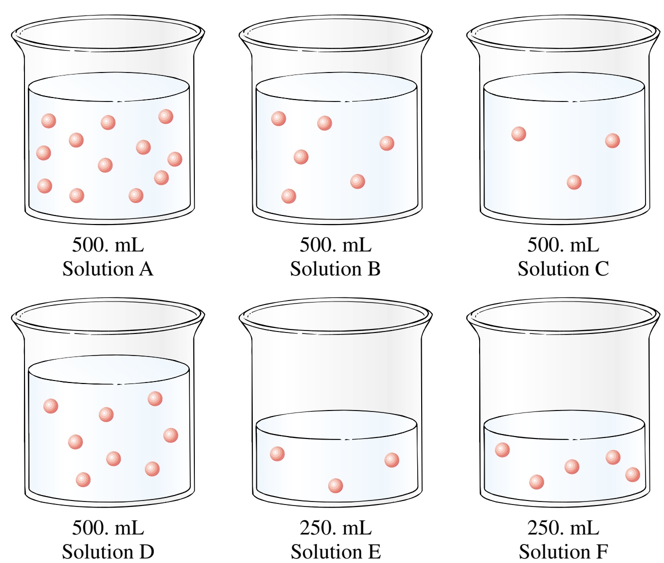
96. A properly written and balanced chemical reaction is critical for the purposes of accurate communication. The oxidation of beryllium is proposed to follow the equation

$$Be + O \longrightarrow BeO$$

(a) Identify the basic error in the writing of this equation.
(b) Provide the balanced equation.

97. How would you prepare 1 L of 1.25×10^{-6} *M* NaCl (molecular weight = 58.44 g/mol) solution by using a balance that can measure mass only to 0.01 g?

98. The drawings shown below represent beakers of aqueous solutions. Each sphere represents a dissolved solute particle.
(a) Which solution is most concentrated?
(b) Which solution is least concentrated?
(c) Which two solutions have the same concentration?
(d) When solutions E and F are combined, the resulting solution has the same concentration as solution ______.



99. You prepared a NaCl solution by adding 58.44 g of NaCl to a 1-L volumetric flask and then adding water to dissolve it. When finished, the final volume in your flask looked like the illustration.

The solution you prepared is
(a) greater than 1 *M* because you added more solvent than necessary.
(b) less than 1 *M* because you added less solvent than necessary.
(c) greater than 1 *M* because you added less solvent than necessary.
(d) less than 1 *M* because you added more solvent than necessary.
(e) is 1 *M* because the amount of solute, not solvent, determines the concentration.

100. Zinc is more active chemically than is silver; it may be used to remove ionic silver from solution.

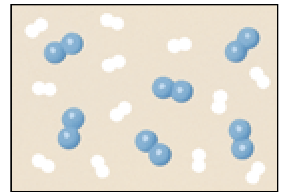
$$Zn(s) + 2AgNO_3(aq) \longrightarrow Zn(NO_3)_2(aq) + 2Ag(s)$$

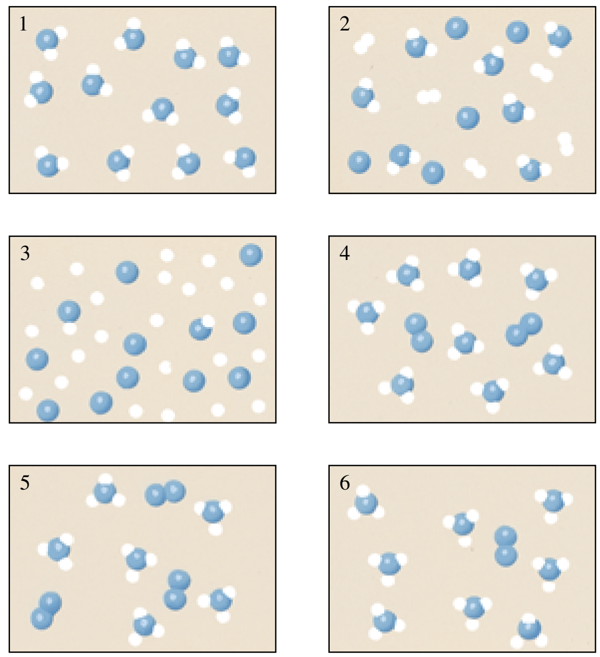
The concentration of a silver nitrate solution is determined to be 1.330 mol/L. Pieces of zinc totaling 100.0 g are added to 1.000 L of the solution; 90.0 g of silver is collected.
(a) Calculate the percent yield of silver.
(b) Suggest a reason why the yield is less than 100.0%.

101. Ammonia is formed in a direct reaction of nitrogen and hydrogen.

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

A tiny portion of the starting mixture is represented by the diagram, where the blue spheres represent N and the white spheres represent H. Which of the following represents the product mixture?





For the reaction of the given sample, which of the following is true?
(a) N_2 is the limiting reactant.
(b) H_2 is the limiting reactant.
(c) NH_3 is the limiting reactant.
(d) No reactant is limiting; they are present in the correct stoichiometric ratio.

BUILDING YOUR KNOWLEDGE

102. Acetic acid, CH_3COOH, reacts with ethanol, CH_3CH_2OH, to form ethyl acetate, $CH_3COOCH_2CH_3$, (density = 0.902 g/mL) by the following reaction.

$$CH_3COOH + CH_3CH_2OH \longrightarrow CH_3COOCH_2CH_3 + H_2O$$

We combine 20.2 mL of acetic acid with 20.1 mL of ethanol.
(a) Which compound is the limiting reactant?
(b) If 27.5 mL of pure ethyl acetate is produced, what is the percent yield? [*Hint:* See Tables 1-1 and 1-8.]

103. Concentrated hydrochloric acid solution is 37.0% HCl and has a density of 1.19 g/mL. A dilute solution of HCl is prepared by diluting 4.50 mL of this concentrated HCl solution to 100.00 mL with water. Then 10.0 mL of this dilute HCl solution reacts with an $AgNO_3$ solution according to the following reaction.

$$HCl(aq) + AgNO_3(aq) \longrightarrow HNO_3(aq) + AgCl(s)$$

How many milliliters of 0.108 *M* $AgNO_3$ solution is required to precipitate all of the chloride as AgCl(s)?

104. In a particular experiment, 272 g of phosphorus, P_4, reacted with excess oxygen to form tetraphosphorus decoxide, P_4O_{10}, in 89.5% yield. In the second step reaction, a 97.8% yield of H_3PO_4 was obtained.
(a) Write the balanced equations for these two reaction steps.
(b) What mass of H_3PO_4 was obtained?

105. Magnesium displaces copper from a dilute solution of copper(II) sulfate; the pure copper will settle out of the solution.

$$Mg(s) + CuSO_4(aq) \longrightarrow MgSO_4(aq) + Cu(s)$$

A copper(II) sulfate solution is mixed by dissolving 25.000 g of copper(II) sulfate, and then it is treated with an excess of magnesium metal. The mass of copper collected is 8.786 g after drying. Calculate the percent yield of copper.

106. Suppose you are designing an experiment for the preparation of hydrogen. For the production of equal amounts of hydrogen, which metal, Zn or Al, is less expensive if Zn costs about half as much as Al on a mass basis?

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$
$$2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$$

107. Gaseous chlorine and gaseous fluorine undergo a combination reaction to form the interhalogen compound ClF.
(a) Write the chemical equation for this reaction.
(b) Calculate the mass of fluorine needed to react with 3.47 g of Cl_2.
(c) How many grams of ClF are formed?

108. It has been estimated that 93% of all atoms in the entire universe are hydrogen and that the vast majority of those remaining are helium. Based on only these two elements and the above-mentioned values, estimate the mass percentage composition of the universe.