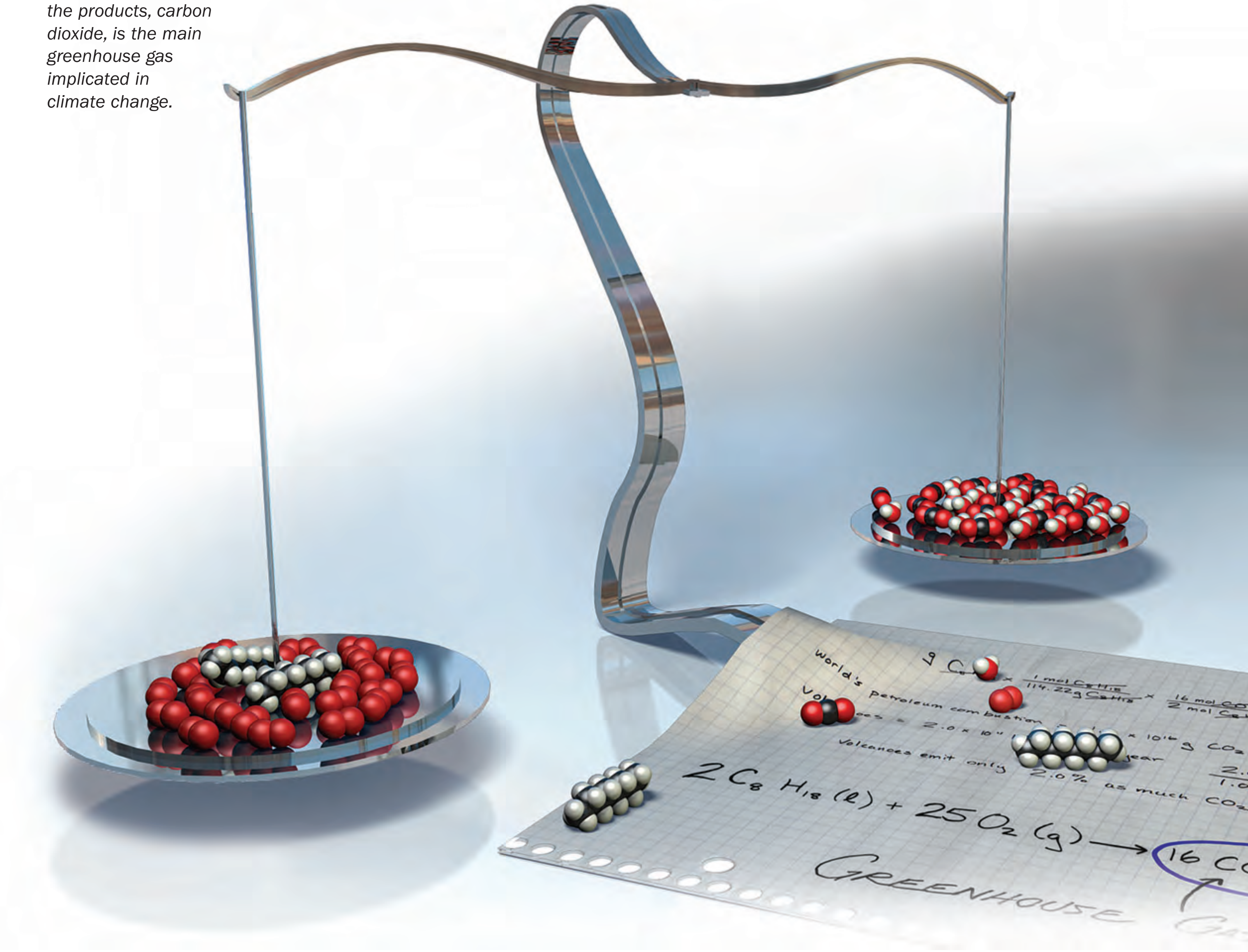
4 Chemical Quantities and Aqueous Reactions



The molecular models on this balance represent the reactants and products in the combustion of octane, a component of petroleum. One of the products, carbon dioxide, is the main greenhouse gas implicated in climate change.

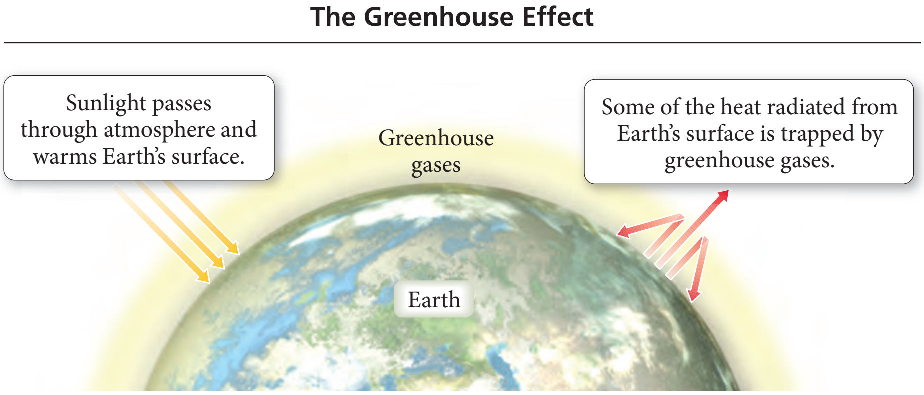
I feel sorry for people who don't understand anything about chemistry. They are missing an important source of happiness.

—Linus Pauling (1901–1994)

WHAT ARE THE RELATIONSHIPS between the amounts of reactants in a chemical reaction and the amounts of products that form? How do we best describe and understand these relationships? The first half of this chapter focuses on chemical stoichiometry—the numerical relationships between the amounts of reactants and products in chemical reactions. In Chapter 3, we wrote balanced chemical equations for chemical reactions. Now we examine more closely the meaning of those balanced equations. In the second half of this chapter, we turn to describing chemical reactions that occur in water. You have probably witnessed many of these types of reactions in your daily life because they are so common. Have you ever mixed baking soda with vinegar and observed the subsequent bubbling? Or have you ever noticed the hard water deposits that form on plumbing fixtures? These reactions—and many others, including those that occur within the watery environment of living cells—are aqueous chemical reactions, the subject of the second half of this chapter.

4.1 Climate Change and the Combustion of Fossil Fuels

The temperature outside my office today is a cool 48 °F, lower than normal for this time of year on the West Coast. However, today's "chill" pales in comparison with how cold it would be without the presence of *greenhouse gases* in the atmosphere. These gases act



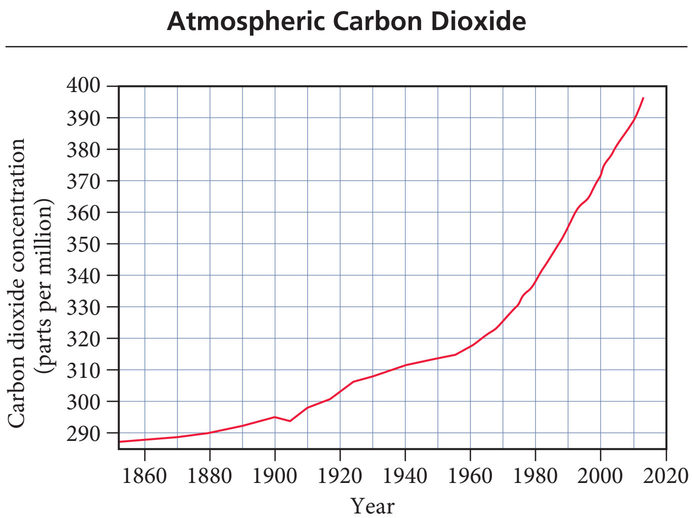
▶ **FIGURE 4.1 The Greenhouse Effect** Greenhouse gases in the atmosphere act as a one-way filter. They allow solar energy to pass through and warm Earth's surface, but they prevent heat from being radiated back out into space.

like the glass of a greenhouse, allowing sunlight to enter the atmosphere and warm Earth's surface, but preventing some of the heat generated by the sunlight from escaping, as shown in FIGURE 4.1▲. The balance between incoming and outgoing energy from the sun determines Earth's average temperature.

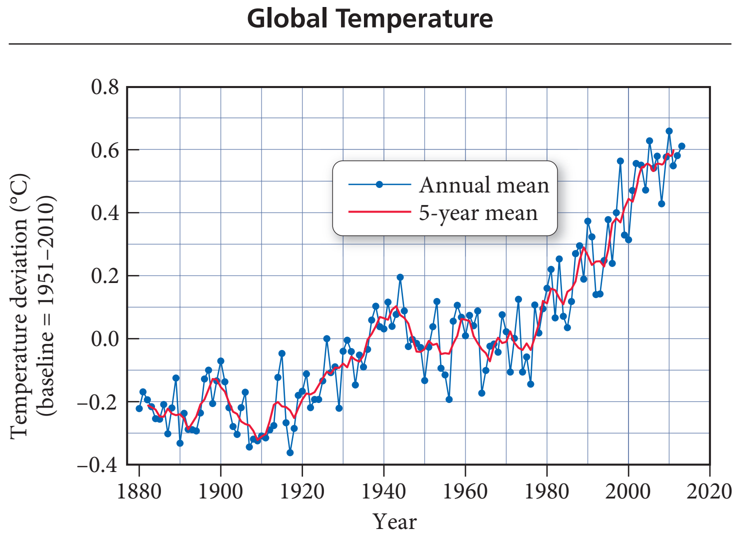
Without greenhouse gases in the atmosphere, more heat energy would escape, and Earth's average temperature would be about 60 °F colder than it is now. The temperature outside of my office today would be below 0 °F, and even the sunniest U.S. cities would most likely be covered with snow. However, if the concentration of greenhouse gases in the atmosphere were to increase, Earth's average temperature would rise.

In recent years scientists have become increasingly concerned because the amount of atmospheric carbon dioxide (CO_2)—Earth's most significant greenhouse gas in terms of climate change—is rising. More CO_2 enhances the atmosphere's ability to hold heat and may therefore lead to *global warming*, an increase in Earth's average temperature. Since 1860, atmospheric CO_2 levels have risen by 38% (FIGURE 4.2▼), and Earth's average temperature has risen by 0.8 °C (about 1.4 °F), as shown in FIGURE 4.3▶.

Most scientists now believe that the primary cause of rising atmospheric CO_2 concentration is the burning of fossil fuels (natural gas, petroleum, and coal), which provides 90% of our society's energy. Some people, however, have suggested that fossil fuel combustion does not significantly contribute to global warming and climate change. They argue, for example, that the amount of carbon dioxide emitted into the atmosphere by volcanic eruptions far exceeds that from fossil fuel combustion. Which group is right? To judge the validity of the naysayers' argument, we need to calculate the amount of carbon dioxide emitted by fossil fuel combustion and compare it to that released by volcanic eruptions. As you will see in the next section of the chapter, at this point in your study of chemistry, you have enough knowledge to do just that.



▶ **FIGURE 4.2 Carbon Dioxide Concentrations in the Atmosphere** The rise in carbon dioxide levels shown in this graph is due largely to fossil fuel combustion.



◀ FIGURE 4.3 **Global Temperature** Average temperatures worldwide have risen by about 0.8 °C since 1880.

4.2 Reaction Stoichiometry: How Much Carbon Dioxide?

KEY CONCEPT VIDEO Reaction Stoichiometry

The amount of carbon dioxide emitted by fossil fuel combustion is related to the amount of fossil fuel that is burned—the balanced chemical equations for the combustion reactions give the exact relationships between these amounts. In this discussion, we use octane (a component of gasoline) as a representative fossil fuel. The balanced equation for the combustion of octane is:

$$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g)$$

The balanced chemical equation shows that 16 CO_2 molecules are produced for every 2 molecules of octane burned. We can extend this numerical relationship between molecules to the amounts in moles as follows:

The coefficients in a chemical reaction specify the relative amounts in moles of each of the substances involved in the reaction.

In other words, from the equation we know that 16 moles of CO_2 are produced for every 2 moles of octane burned. The numerical relationship between chemical amounts in a balanced chemical equation is reaction **stoichiometry**. Stoichiometry allows us to predict the amounts of products that will form in a chemical reaction based on the amounts of reactants that undergo the reaction. Stoichiometry also allows us to determine the amount of reactants necessary to form a given amount of product. These calculations are central to chemistry, allowing chemists to plan and carry out chemical reactions to obtain products in the desired quantities.

Stoichiometry is pronounced stoy-kee-AHM-e-tree.

Making Pizza: The Relationships Among Ingredients

The concepts of stoichiometry are similar to those in a cooking recipe. Calculating the amount of carbon dioxide produced by the combustion of a given amount of a fossil fuel is analogous to calculating the number of pizzas that can be made with a given amount of cheese. For example, suppose we use the following pizza recipe:

$$1 \text{ crust} + 5 \text{ ounces tomato sauce} + 2 \text{ cups cheese} \longrightarrow 1 \text{ pizza}$$

The recipe shows the numerical relationships between the pizza ingredients. It says that if we have 2 cups of cheese—and enough of everything else—we can make 1 pizza. We can write this relationship as a ratio between the cheese and the pizza.

$$2 \text{ cups cheese} : 1 \text{ pizza}$$

What if we have 6 cups of cheese? Assuming that we have enough of everything else, we can use the above ratio as a conversion factor to calculate the number of pizzas.

$$6\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 3\ \text{pizzas}$$

Six cups of cheese are sufficient to make 3 pizzas. The pizza recipe contains numerical ratios between other ingredients as well, including the following:

$$1\ \text{crust} : 1\ \text{pizza}$$
$$5\ \text{ounces tomato sauce} : 1\ \text{pizza}$$

Making Molecules: Mole-to-Mole Conversions

A balanced chemical equation is a "recipe" for how reactants combine to form products. From our balanced equation for the combustion of octane, for example, we can write the following stoichiometric ratio:

$$2\ \text{mol}\ C_8H_{18}(l) : 16\ \text{mol}\ CO_2(g)$$

We can use this ratio to determine how many moles of CO_2 are produced for a given number of moles of C_8H_{18} burned. Suppose that we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 are produced? We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe. This ratio acts as a conversion factor allowing us to convert from the amount in moles of the reactant (C_8H_{18}) to the amount in moles of the product (CO_2):

$$22.0\ \cancel{\text{mol}\ C_8H_{18}} \times \frac{16\ \text{mol}\ CO_2}{2\ \cancel{\text{mol}\ C_8H_{18}}} = 176\ \text{mol}\ CO_2$$

The combustion of 22 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

Making Molecules: Mass-to-Mass Conversions

According to the U.S. Department of Energy, the world burned 3.3×10^{10} barrels of petroleum in 2013, the equivalent of approximately 3.7×10^{15} g of gasoline. We can estimate the mass of CO_2 emitted into the atmosphere from burning this much gasoline by using the combustion of 3.7×10^{15} g octane as the representative reaction. This calculation is similar to the one we just did, except that we are given the *mass* of octane instead of the *amount* of octane in moles. Consequently, we must first convert the mass (in grams) to the amount (in moles). The general conceptual plan for calculations in which we are given the mass of a reactant or product in a chemical reaction and asked to find the mass of a different reactant or product is

$$\boxed{\text{Mass A}} \longrightarrow \boxed{\begin{array}{c}\text{Amount A}\\ \text{(in moles)}\end{array}} \longrightarrow \boxed{\begin{array}{c}\text{Amount B}\\ \text{(in moles)}\end{array}} \longrightarrow \boxed{\text{Mass B}}$$

where A and B are two different substances involved in the reaction. We use the molar mass of A to convert from the mass of A to the amount of A (in moles). We use the appropriate ratio from the balanced chemical equation to convert from the amount of A (in moles) to the amount of B (in moles). And finally, we use the molar mass of B to convert from the amount of B (in moles) to the mass of B. To calculate the mass of CO_2 emitted upon the combustion of 3.7×10^{15} g of octane, therefore, we use the following conceptual plan:

Conceptual Plan

$$\boxed{\text{g}\ C_8H_{18}} \xrightarrow[\frac{1\ \text{mol}\ C_8H_{18}}{114.22\ \text{g}\ C_8H_{18}}]{} \boxed{\text{mol}\ C_8H_{18}} \xrightarrow[\frac{16\ \text{mol}\ CO_2}{2\ \text{mol}\ C_8H_{18}}]{} \boxed{\text{mol}\ CO_2} \xrightarrow[\frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2}]{} \boxed{\text{g}\ CO_2}$$

Relationships Used

$2\ \text{mol}\ C_8H_{18} : 16\ \text{mol}\ CO_2$ (from the chemical equation)
molar mass $C_8H_{18} = 114.22\ \text{g/mol}$
molar mass $CO_2 = 44.01\ \text{g/mol}$

Solution

We follow the conceptual plan to solve the problem, beginning with g C_8H_{18} and canceling units to arrive at g CO_2:

$$3.7 \times 10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.22\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\ \text{g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.1 \times 10^{16}\ \text{g } CO_2$$

The world's petroleum combustion produces 1.1×10^{16} g CO_2 (1.1×10^{13} kg) per year. In comparison, volcanoes produce about 2.0×10^{11} kg CO_2 per year.* In other words, volcanoes emit only $\frac{2.0 \times 10^{11}\ \text{kg}}{1.1 \times 10^{13}\ \text{kg}} \times 100\% = 1.8\%$ as much CO_2 per year as petroleum combustion. The argument that volcanoes emit more carbon dioxide than fossil fuel combustion is blatantly incorrect. Examples 4.1 and 4.2 provide additional practice with stoichiometric calculations.

The percentage of CO_2 emitted by volcanoes relative to all fossil fuels is even less than 2% because CO_2 is also emitted by the combustion of coal and natural gas.

EXAMPLE 4.1 Stoichiometry

During photosynthesis, plants convert carbon dioxide and water into glucose ($C_6H_{12}O_6$) according to the reaction:

$$6\ CO_2(g) + 6\ H_2O(l) \xrightarrow{\text{sunlight}} 6\ O_2(g) + C_6H_{12}O_6(aq)$$

Suppose a particular plant consumes 37.8 g CO_2 in one week. Assuming that there is more than enough water present to react with all of the CO_2, what mass of glucose (in grams) can the plant synthesize from the CO_2?

SORT The problem gives the mass of carbon dioxide and asks you to find the mass of glucose that can be produced.

GIVEN 37.8 g CO_2
FIND g $C_6H_{12}O_6$

STRATEGIZE The conceptual plan follows the general pattern of mass A $\rightarrow$ amount A (in moles) $\rightarrow$ amount B (in moles) $\rightarrow$ mass B. From the chemical equation, you can deduce the relationship between moles of carbon dioxide and moles of glucose. Use the molar masses to convert between grams and moles.

CONCEPTUAL PLAN

$$\boxed{\text{g } CO_2} \xrightarrow[\frac{1\ \text{mol } CO_2}{44.01\ \text{g } CO_2}]{} \boxed{\text{mol } CO_2} \xrightarrow[\frac{1\ \text{mol } C_6H_{12}O_6}{6\ \text{mol } CO_2}]{} \boxed{\text{mol } C_6H_{12}O_6} \xrightarrow[\frac{180.16\ \text{g } C_6H_{12}O_6}{1\ \text{mol } C_6H_{12}O_6}]{} \boxed{\text{g } C_6H_{12}O_6}$$

RELATIONSHIPS USED

molar mass $CO_2 = 44.01$ g/mol
6 mol CO_2 : 1 mol $C_6H_{12}O_6$
molar mass $C_6H_{12}O_6 = 180.16$ g/mol

SOLVE Follow the conceptual plan to solve the problem. Begin with g CO_2 and use the conversion factors to arrive at g $C_6H_{12}O_6$.

SOLUTION

$$37.8\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mol } CO_2}}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{6\ \cancel{\text{mol } CO_2}} \times \frac{180.16\ \text{g } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 25.8\ \text{g } C_6H_{12}O_6$$

CHECK The units of the answer are correct. The magnitude of the answer (25.8 g) is less than the initial mass of CO_2 (37.8 g). This is reasonable because each carbon in CO_2 has two oxygen atoms associated with it, while in $C_6H_{12}O_6$ each carbon has only one oxygen atom associated with it and two hydrogen atoms, which are much lighter than oxygen. Therefore the mass of glucose produced should be less than the mass of carbon dioxide for this reaction.

FOR PRACTICE 4.1

Magnesium hydroxide, the active ingredient in milk of magnesia, neutralizes stomach acid, primarily HCl, according to the reaction:

$$Mg(OH)_2(aq) + 2\ HCl(aq) \rightarrow 2\ H_2O(l) + MgCl_2(aq)$$

What mass of HCl, in grams, is neutralized by a dose of milk of magnesia containing 3.26 g $Mg(OH)_2$?

*Gerlach, T. M., Present-day CO_2 emissions from volcanoes; *Eos, Transactions, American Geophysical Union*, Vol. 72, No. 23, June 4, 1991, pp. 249 and 254–255

EXAMPLE 4.2 Stoichiometry

Sulfuric acid (H_2SO_4) is a component of acid rain that forms when SO_2, a pollutant, reacts with oxygen and water according to the simplified reaction:

$$2\ SO_2(g) + O_2(g) + 2\ H_2O(l) \longrightarrow 2\ H_2SO_4(aq)$$

The generation of the electricity used in a typical medium-sized home produces about 25 kg of SO_2 per year. Assuming that there is more than enough O_2 and H_2O, what mass of H_2SO_4, in kg, can form from this much SO_2?

SORT The problem gives the mass of sulfur dioxide and asks you to find the mass of sulfuric acid.

GIVEN 25 kg SO_2
FIND kg H_2SO_4

STRATEGIZE The conceptual plan follows the standard format of mass $\longrightarrow$ amount (in moles) $\longrightarrow$ amount (in moles) $\longrightarrow$ mass. Because the original quantity of SO_2 is given in kg, you must first convert to grams. You can deduce the relationship between moles of sulfur dioxide and moles of sulfuric acid from the chemical equation. Since the final quantity is requested in kg, convert to kg at the end.

CONCEPTUAL PLAN

$$\text{kg } SO_2 \xrightarrow{\frac{1000\text{ g}}{1\text{ kg}}} \text{g } SO_2 \xrightarrow{\frac{1\text{ mol } SO_2}{64.06\text{ g } SO_2}} \text{mol } SO_2 \xrightarrow{\frac{2\text{ mol } H_2SO_4}{2\text{ mol } SO_2}}$$

$$\text{mol } H_2SO_4 \xrightarrow{\frac{98.08\text{ g } H_2SO_4}{1\text{ mol } H_2SO_4}} \text{g } H_2SO_4 \xrightarrow{\frac{1\text{ kg}}{1000\text{ g}}} \text{kg } H_2SO_4$$

RELATIONSHIPS USED
1 kg = 1000 g
molar mass SO_2 = 64.06 g/mol
2 mol SO_2 : 2 mol H_2SO_4
molar mass H_2SO_4 = 98.08 g/mol

SOLVE Follow the conceptual plan to solve the problem. Begin with the given amount of SO_2 in kilograms and use the conversion factors to arrive at kg H_2SO_4.

SOLUTION

$$25\ \cancel{\text{kg}}\ \cancel{SO_2} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol } SO_2}}{64.06\ \cancel{\text{g}}\ \cancel{SO_2}} \times \frac{2\ \cancel{\text{mol } H_2SO_4}}{2\ \cancel{\text{mol } SO_2}}$$

$$\times \frac{98.08\ \cancel{\text{g}}\ H_2SO_4}{1\ \cancel{\text{mol } H_2SO_4}} \times \frac{1\text{ kg}}{1000\ \cancel{\text{g}}} = 38\text{ kg } H_2SO_4$$

CHECK The units of the final answer are correct. The magnitude of the final answer (38 kg H_2SO_4) is larger than the amount of SO_2 given (25 kg). This is reasonable because in the reaction each SO_2 molecule "gains weight" by reacting with O_2 and H_2O.

FOR PRACTICE 4.2

Another component of acid rain is nitric acid, which forms when NO_2, also a pollutant, reacts with oxygen and water according to the simplified equation:

$$4\ NO_2(g) + O_2(g) + 2\ H_2O(l) \longrightarrow 4\ HNO_3(aq)$$

The generation of the electricity used in a medium-sized home produces about 16 kg of NO_2 per year. Assuming that there is plenty of O_2 and H_2O, what mass of HNO_3, in kg, can form from this amount of NO_2 pollutant?

STOICHIOMETRY I

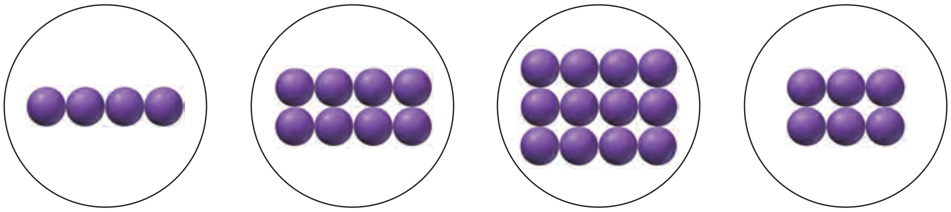
CONCEPTUAL CONNECTION 4.1

Under certain conditions, sodium can react with oxygen to form sodium oxide according to the reaction:

$$4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$$

A flask contains the amount of oxygen represented by the diagram to the right.

Which image best represents the amount of sodium required to completely react with all of the oxygen in the flask?

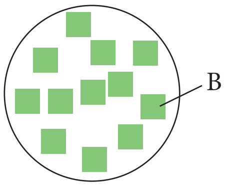


(a) **(b)** **(c)** **(d)**

STOICHIOMETRY II

CONCEPTUAL CONNECTION 4.2

Consider the generic chemical equation A + 3B ⟶ 2C. Let circles represent molecules of A, squares represent molecules of B, and triangles represent molecules of C. The diagram shown here represents the amount of B available for reaction. Draw similar diagrams showing (a) the amount of A necessary to completely react with B, and (b) the amount of C that forms if A completely reacts.



4.3 Limiting Reactant, Theoretical Yield, and Percent Yield

KEY CONCEPT VIDEO
Limiting Reactant, Theoretical Yield, and Percent Yield

Let's return to our pizza analogy to understand three more concepts important in reaction stoichiometry: *limiting reactant*, *theoretical yield*, and *percent yield*. Recall the pizza recipe from Section 4.2:

$$1 \text{ crust} + 5 \text{ ounces tomato sauce} + 2 \text{ cups cheese} \longrightarrow 1 \text{ pizza}$$

Suppose that we have 4 crusts, 10 cups of cheese, and 15 ounces of tomato sauce. How many pizzas can we make?

We have enough crusts to make:

$$4\,\cancel{\text{crusts}} \times \frac{1 \text{ pizza}}{1\,\cancel{\text{crust}}} = 4 \text{ pizzas}$$

We have enough cheese to make:

$$10\,\cancel{\text{cups cheese}} \times \frac{1 \text{ pizza}}{2\,\cancel{\text{cups cheese}}} = 5 \text{ pizzas}$$

We have enough tomato sauce to make:

$$15\ \cancel{\text{ounces tomato sauce}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = 3\ \text{pizzas}$$

Limiting reactant

Smallest number of pizzas

The term *limiting reagent* is sometimes used in place of *limiting reactant*.

We have enough crusts for 4 pizzas, enough cheese for 5 pizzas, but enough tomato sauce for only 3 pizzas. Consequently, unless we get more ingredients, we can make only 3 pizzas. The tomato sauce *limits* how many pizzas we can make. If the pizza recipe were a chemical reaction, the tomato sauce would be the **limiting reactant**, the reactant that limits the amount of product in a chemical reaction. Notice that the limiting reactant is simply the reactant that makes *the least amount of product.* Reactants that *do not* limit the amount of product—such as the crusts and the cheese in this example—are said to be *in excess*. If this were a chemical reaction, 3 pizzas would be the **theoretical yield**, the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.

Let us carry this analogy one step further. Suppose we go on to cook our pizzas and accidentally burn one of them. Even though we theoretically have enough ingredients for 3 pizzas, we end up with only 2. If this were a chemical reaction, the 2 pizzas would be our **actual yield**, the amount of product actually produced by a chemical reaction. (The actual yield is always equal to or less than the theoretical yield because at least a small amount of product is usually lost to other reactions or does not form during a reaction.) Finally, we calculate our **percent yield**, the percentage of the theoretical yield that was actually attained, as follows:

Actual yield

$$\%\ \text{yield} = \frac{2\ \cancel{\text{pizzas}}}{3\ \cancel{\text{pizzas}}} \times 100\% = 67\%$$

Theoretical yield

Since one of our pizzas burned, we obtained only 67% of our theoretical yield.

Summarizing: Limiting Reactant and Yield

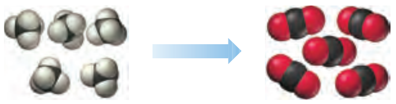
- **The limiting reactant** (or **limiting reagent**) is the reactant that is completely consumed in a chemical reaction and limits the amount of product.
- **The reactant in excess** is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.
- **The theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- **The actual yield** is the amount of product actually produced by a chemical reaction.
- **The percent yield** is calculated as $\dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$.

We can apply these concepts to a chemical reaction. Recall from Section 3.10 our balanced equation for the combustion of methane:

$$CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(l)$$

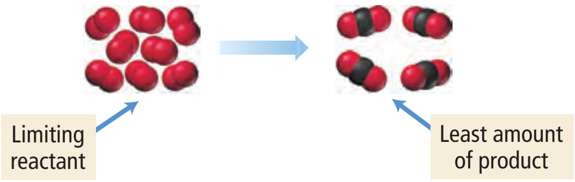
If we start out with 5 CH_4 molecules and 8 O_2 molecules, what is our limiting reactant? What is our theoretical yield of carbon dioxide molecules? We first calculate the number of CO_2 molecules that can be made from 5 CH_4 molecules:

$$5\ \cancel{CH_4} \times \frac{1\ CO_2}{1\ \cancel{CH_4}} = 5\ CO_2$$



Next, we calculate the number of CO_2 molecules that can be made from 8 O_2 molecules:

$$8\ \cancel{O_2} \times \frac{1\ CO_2}{2\ \cancel{O_2}} = 4\ CO_2$$



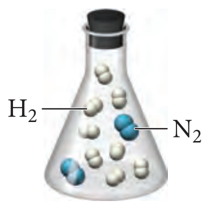
We have enough CH_4 to make 5 CO_2 molecules and enough O_2 to make 4 CO_2 molecules; therefore O_2 is the limiting reactant and 4 CO_2 molecules is the theoretical yield. The CH_4 is in excess.

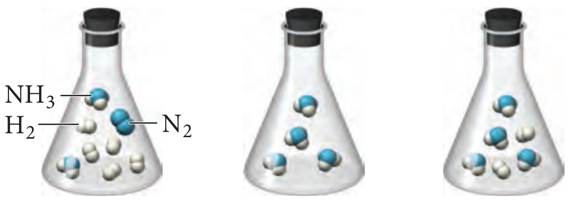
LIMITING REACTANT AND THEORETICAL YIELD — **CONCEPTUAL CONNECTION 4.3**

Nitrogen and hydrogen gas react to form ammonia according to the reaction:

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

If a flask contains a mixture of reactants represented by the image shown at right, which of the following images best represents the mixture after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?





(a) (b) (c)

Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Reactant Masses

When working in the laboratory, we normally measure the initial quantities of reactants in grams, not in number of molecules. To find the limiting reactant and theoretical yield from initial masses, we must first convert the masses to amounts in moles. Consider, for example, the following reaction:

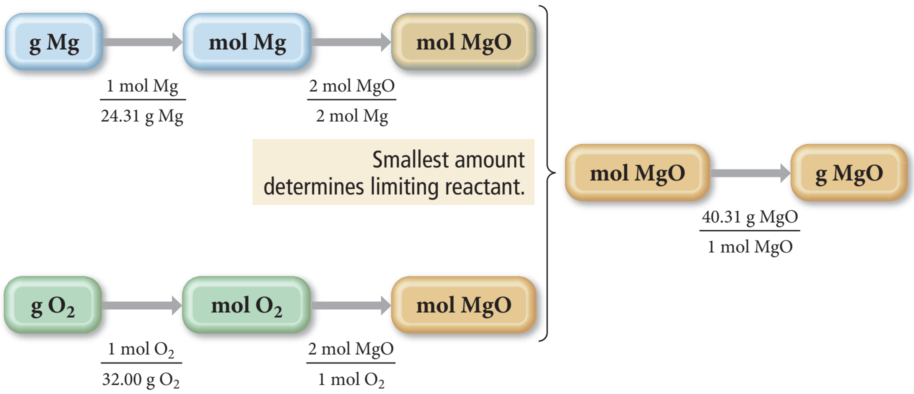
$$2\ Mg(s) + O_2(g) \longrightarrow 2\ MgO(s)$$

If we have 42.5 g Mg and 33.8 g O_2, what is the limiting reactant and theoretical yield?

To solve this problem, we must determine which of the reactants makes the least amount of product.

Conceptual Plan

We find the limiting reactant by calculating how much product can be made from each reactant. Since we are given the initial quantities in grams, and stoichiometric relationships are between moles, we must first convert to moles. We then convert from moles of the reactant to moles of product. The reactant that makes the *least amount of product* is the limiting reactant. The conceptual plan is:



In this conceptual plan, we compare the number of moles of MgO made by each reactant and convert only the smaller amount to grams. (Alternatively, we can convert both quantities to grams and determine the limiting reactant based on the mass of the product.)

Relationships Used

molar mass Mg $= 24.31$ g/mol
molar mass $O_2 = 32.00$ g/mol
2 mol Mg : 2 mol MgO
1 mol O_2 : 2 mol MgO
molar mass MgO $= 40.31$ g/mol

Solution

Beginning with the masses of each reactant, we follow the conceptual plan to calculate how much product can be made from each.

$$\boxed{42.5\ \cancel{\text{g Mg}}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = \boxed{1.7\underline{4}83\ \text{mol MgO}}$$

Limiting reactant

Least amount of product

$$33.8\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol O}_2}} = 2.1\underline{1}25\ \text{mol MgO}$$

$$1.7\underline{4}83\ \cancel{\text{mol MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = \mathbf{70.5\ g\ MgO}$$

Since Mg makes the least amount of product, it is the limiting reactant and O_2 is in excess. Notice that the limiting reactant is not necessarily the reactant with the least mass. In this case, the mass of O_2 is less than the mass of Mg, yet Mg is the limiting reactant because it makes the least amount of MgO. The theoretical yield is therefore 70.5 g of MgO, the mass of product possible based on the limiting reactant.

Now suppose that when the synthesis is carried out, the actual yield of MgO is 55.9 g. What is the percent yield? We calculate the percent yield as follows:

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{55.9\ \cancel{\text{g}}}{70.5\ \cancel{\text{g}}} \times 100\% = 79.3\%$$

EXAMPLE 4.3 Limiting Reactant and Theoretical Yield

Ammonia, NH_3, can be synthesized by the reaction:

$$2\ NO(g) + 5\ H_2(g) \longrightarrow 2\ NH_3(g) + 2\ H_2O(g)$$

Starting with 86.3 g NO and 25.6 g H_2, find the theoretical yield of ammonia in grams.

SORT You are given the mass of each reactant in grams and asked to find the theoretical yield of a product.

GIVEN 86.3 g NO, 25.6 g H_2
FIND theoretical yield of NH_3

STRATEGIZE Determine which reactant makes the least amount of product by converting from grams of each reactant to moles of the reactant to moles of the product. Use molar masses to convert between grams and moles and use the stoichiometric relationships (deduced from the chemical equation) to convert between moles of reactant and moles of product. The reactant that makes *the least amount of product* is the limiting reactant. Convert the number of moles of product—obtained using the limiting reactant—to grams of product.

CONCEPTUAL PLAN

$$\text{g NO} \xrightarrow{\frac{1\text{ mol NO}}{30.01\text{ g NO}}} \text{mol NO} \xrightarrow{\frac{2\text{ mol NH}_3}{2\text{ mol NO}}} \text{mol NH}_3$$

$$\text{g H}_2 \xrightarrow{\frac{1\text{ mol H}_2}{2.016\text{ g H}_2}} \text{mol H}_2 \xrightarrow{\frac{2\text{ mol NH}_3}{5\text{ mol H}_2}} \text{mol NH}_3$$

Smallest amount determines limiting reactant.

$$\text{mol NH}_3 \xrightarrow{\frac{17.03\text{ g NH}_3}{1\text{ mol NH}_3}} \text{g NH}_3$$

RELATIONSHIPS USED
molar mass NO = 30.01 g/mol
molar mass H_2 = 2.016 g/mol
2 mol NO : 2 mol NH_3 (from chemical equation)
5 mol H_2 : 2 mol NH_3 (from chemical equation)
molar mass NH_3 = 17.03 g/mol

SOLVE Start with the given mass of each reactant and calculate the amount of product, in moles, that can be made. Convert the amount of product made by the limiting reactant to grams—this is the theoretical yield.

SOLUTION

89%

$$\mathbf{86.3\ \cancel{g\ NO}} \times \frac{1\ \cancel{\text{mol NO}}}{30.01\ \cancel{\text{g NO}}} \times \frac{2\text{ mol NH}_3}{2\ \cancel{\text{mol NO}}} = \mathbf{2.87\underline{5}7\ mol\ NH_3}$$

Limiting reactant

Least amount of product

$$\mathbf{25.6\ \cancel{g\ H_2}} \times \frac{1\ \cancel{\text{mol H}_2}}{2.016\ \cancel{\text{g H}_2}} \times \frac{2\text{ mol NH}_3}{5\ \cancel{\text{mol H}_2}} = \mathbf{5.0\underline{7}94\ mol\ NH_3}$$

$$\mathbf{2.87\underline{5}7\ \cancel{mol\ NH_3}} \times \frac{17.03\text{ g NH}_3}{\cancel{\text{mol NH}_3}} = \mathbf{49.0\ g\ NH_3}$$

Since NO makes the least amount of product, it is the limiting reactant, and the theoretical yield of ammonia is 49.0 g.

—*Continued on the next page*

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CHECK The units of the answer (g NH_3) are correct. The magnitude (49.0 g) seems reasonable given that 86.3 g NO is the limiting reactant. NO contains one oxygen atom per nitrogen atom, and NH_3 contains 3 hydrogen atoms per nitrogen atom. Since 3 hydrogen atoms have less mass than 1 oxygen atom, it is reasonable that the mass of NH_3 obtained is less than the mass of NO.

FOR PRACTICE 4.3

Ammonia can also be synthesized by the reaction:

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

What is the theoretical yield of ammonia, in kg, that can be synthesized from 5.22 kg of H_2 and 31.5 kg of N_2?

EXAMPLE 4.4 Limiting Reactant and Theoretical Yield

Titanium metal can be obtained from its oxide according to the balanced equation:

$$TiO_2(s) + 2\ C(s) \longrightarrow Ti(s) + 2\ CO(g)$$

When 28.6 kg of C reacts with 88.2 kg of TiO_2, 42.8 kg of Ti are produced. Determine the limiting reactant, theoretical yield (in kg), and percent yield.

SORT You are given the mass of each reactant and the mass of product formed. You are asked to find the limiting reactant, theoretical yield, and percent yield.

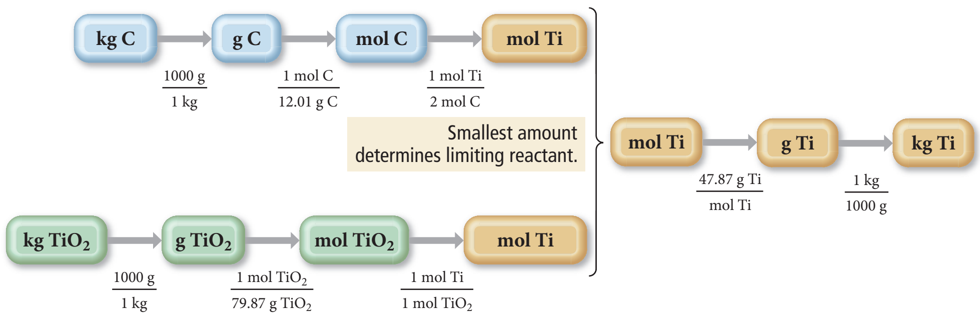
GIVEN 28.6 kg C, 88.2 kg TiO_2, 42.8 kg Ti produced

FIND limiting reactant, theoretical yield, % yield

STRATEGIZE Determine which of the reactants makes the least amount of product by converting from kilograms of each reactant to moles of product. Convert between grams and moles using molar mass. Convert between moles of reactant and moles of product using the stoichiometric relationships derived from the chemical equation. The reactant that makes the *least amount of product* is the limiting reactant.

Determine the theoretical yield (in kg) by converting the number of moles of product—obtained with the limiting reactant—to kilograms of product.

CONCEPTUAL PLAN



RELATIONSHIPS USED

1000 g = 1 kg

molar mass of C = 12.01 g/mol

molar mass of TiO_2 = 79.87 g/mol

1 mol TiO_2 : 1 mol Ti

2 mol C : 1 mol Ti

molar mass of Ti = 47.87 g/mol

SOLVE Beginning with the actual amount of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to kilograms—this is the theoretical yield.	**SOLUTION** $$\mathbf{28.6\ \cancel{kg\ C}} \times \frac{1000\ \cancel{g}}{1\ \cancel{kg}} \times \frac{1\ \cancel{mol\ C}}{12.01\ \cancel{g\ C}} \times \frac{1\ \text{mol Ti}}{2\ \cancel{mol\ C}} = \mathbf{1.1907 \times 10^3\ mol\ Ti}$$ Limiting reactant; Least amount of product $$\boxed{\mathbf{88.2\ \cancel{kg\ TiO_2}}} \times \frac{1000\ \cancel{g}}{1\ \cancel{kg}} \times \frac{1\ \cancel{mol\ TiO_2}}{79.87\ \cancel{g\ TiO_2}} \times \frac{1\ \cancel{mol\ Ti}}{1\ \cancel{mol\ TiO_2}} = \boxed{\mathbf{1.1043 \times 10^3\ mol\ Ti}}$$ $$\mathbf{1.1043 \times 10^3\ \cancel{mol\ Ti}} \times \frac{47.87\ \cancel{g\ Ti}}{1\ \cancel{mol\ Ti}} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} = \mathbf{52.9\ kg\ Ti}$$ Since TiO_2 makes the least amount of product, it is the limiting reactant and 52.9 kg Ti is the theoretical yield.
Calculate the percent yield by dividing the actual yield (42.8 kg Ti) by the theoretical yield.	$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{42.8\ \cancel{kg}}{52.9\ \cancel{kg}} \times 100 = 80.9\%$$

CHECK The theoretical yield has the correct units (kg Ti) and has a reasonable magnitude compared to the mass of TiO_2. Since Ti has a lower molar mass than TiO_2, the amount of Ti made from TiO_2 should have a lower mass. The percent yield is reasonable (under 100% as it should be).

FOR PRACTICE 4.4

The following reaction is used to obtain iron from iron ore:

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

The reaction of 167 g Fe_2O_3 with 85.8 g CO produces 72.3 g Fe. Determine the limiting reactant, theoretical yield, and percent yield.

REACTANT IN EXCESS — CONCEPTUAL CONNECTION 4.4

Nitrogen dioxide reacts with water to form nitric acid and nitrogen monoxide according to the equation:

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

Suppose that 5 mol NO_2 and 1 H_2O mol combine and react completely. How many moles of the reactant in excess are present after the reaction has completed?

4.4 Solution Concentration and Solution Stoichiometry

Chemical reactions in which the reactants are dissolved in water are among the most common and important. The reactions that occur in lakes, streams, and oceans, as well as the reactions that occur in every cell of our bodies, take place in water. A homogeneous mixture of two or more substances—such as salt and water—is a **solution**. The majority component of a solution is the **solvent**, and the minority component is the **solute**. A solution in which water acts as the solvent is an **aqueous solution**. In this section, we first examine how to quantify the concentration of a solution (the amount of solute relative to solvent) and then turn to applying the principles of stoichiometry, which we discussed in the previous section, to reactions occurring in solution.

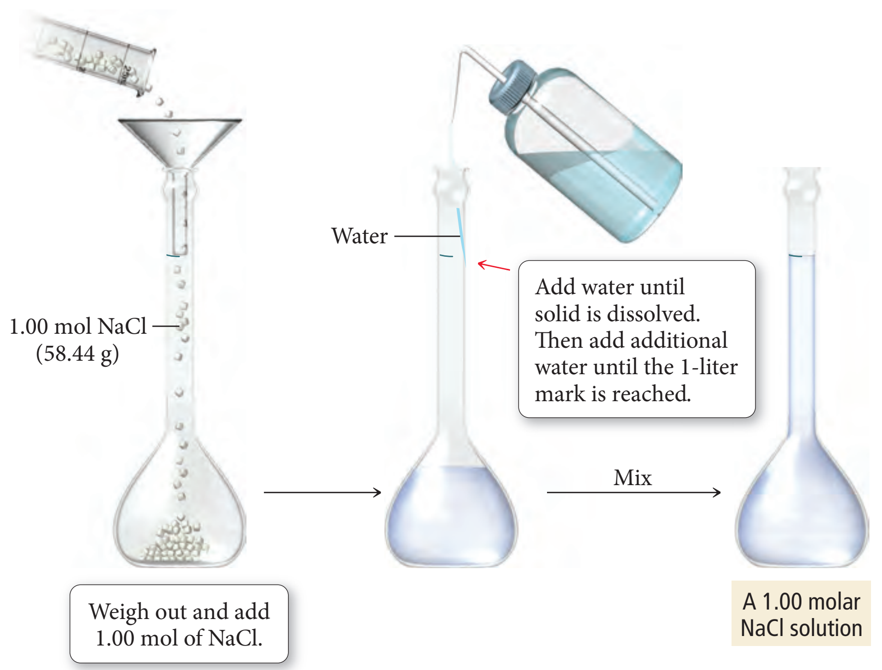
Solution Concentration

The amount of solute in a solution is variable. For example, we can add just a little salt to water to make a **dilute solution**, one that contains a small amount of solute relative to the solvent, or we can add a lot of salt to water to make a **concentrated solution**, one that contains a large amount of solute relative to the solvent. A common way to express solution concentration is **molarity (M)**, the amount of solute (in moles) divided by the volume of solution (in liters).

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

Note that molarity is a ratio of the amount of solute per liter of *solution*, not per liter of solvent. To make an aqueous solution of a specified molarity, we usually put the solute into a flask and then add water until we have the desired volume of solution. For example, to make 1 L of a 1 M NaCl solution, we add 1 mol of NaCl to a flask and then add water to make 1 L of solution (FIGURE 4.4▼). We *do not* combine 1 mol of NaCl with 1 L of water because the resulting solution would have a total volume exceeding 1 L and therefore a molarity of less than 1 M. To calculate molarity, we divide the amount of the solute in moles by the volume of the solution (solute *and* solvent) in liters, as shown in Example 4.5.

Preparing a Solution of Specified Concentration



▲ FIGURE 4.4 **Preparing a 1.00 Molar NaCl Solution**

EXAMPLE 4.5 Calculating Solution Concentration

What is the molarity of a solution containing 25.5 g KBr dissolved in enough water to make 1.75 L of solution?

SORT You are given the mass of KBr and the volume of a solution and asked to find its molarity.

GIVEN 25.5 g KBr, 1.75 L of solution
FIND molarity (M)

STRATEGIZE When formulating the conceptual plan, think about the definition of molarity, the amount of solute *in moles* per liter of solution.
You are given the mass of KBr, so first use the molar mass of KBr to convert from g KBr to mol KBr.
Then use the number of moles of KBr and liters of solution to find the molarity.

CONCEPTUAL PLAN

$$\text{g KBr} \xrightarrow{\frac{1\ \text{mol}}{119.00\ \text{g}}} \text{mol KBr}$$

$$\text{mol KBr, L solution} \longrightarrow \text{Molarity}$$

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

RELATIONSHIP USED molar mass of KBr = 119.00 g/mol

SOLVE Follow the conceptual plan. Begin with g KBr and convert to mol KBr; then use mol KBr and L solution to calculate molarity.

SOLUTION

$$25.5\ \cancel{\text{g KBr}} \times \frac{1\ \text{mol KBr}}{119.00\ \cancel{\text{g KBr}}} = 0.21\underline{4}29\ \text{mol KBr}$$

$$\text{molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

$$= \frac{0.21\underline{4}29\ \text{mol KBr}}{1.75\ \text{L solution}} = 0.122\ \text{M}$$

CHECK The units of the answer (M) are correct. The magnitude is reasonable. Common solutions range in concentration from 0 to about 18 M. Concentrations significantly above 18 M are suspect and should be double-checked.

FOR PRACTICE 4.5

Calculate the molarity of a solution made by adding 45.4 g of $NaNO_3$ to a flask and dissolving with water to a total volume of 2.50 L.

FOR MORE PRACTICE 4.5

What mass of KBr (in grams) should you use to make 250.0 mL of a 1.50 M KBr solution?

Using Molarity in Calculations

The molarity of a solution can be used as a conversion factor between moles of the solute and liters of the solution. For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.

$$\frac{0.500\ \text{mol NaCl}}{\text{L solution}} \quad \textit{converts} \quad \text{L solution} \longrightarrow \text{mol NaCl}$$

This conversion factor converts from L solution to mol NaCl. If you want to convert in the other direction, invert the conversion factor.

$$\frac{\text{L solution}}{0.500\ \text{mol NaCl}} \quad \textit{converts} \quad \text{mol NaCl} \longrightarrow \text{L solution}$$

Example 4.6 illustrates how to use molarity in this way.

EXAMPLE 4.6 Using Molarity in Calculations

How many liters of a 0.125 M NaOH solution contain 0.255 mol of NaOH?

SORT You are given the concentration of a NaOH solution. You are asked to find the volume of the solution that contains a given amount (in moles) of NaOH.	**GIVEN** 0.125 M NaOH solution, 0.255 mol NaOH **FIND** volume of NaOH solution (in L)
STRATEGIZE The conceptual plan begins with mol NaOH and shows the conversion to L of solution using the molarity as a conversion factor.	**CONCEPTUAL PLAN** $\text{mol NaOH} \xrightarrow{\frac{1\text{ L solution}}{0.125\text{ mol NaOH}}} \text{L solution}$ **RELATIONSHIPS USED** $0.125\text{ M NaOH} = \dfrac{0.125\text{ mol NaOH}}{1\text{ L solution}}$
SOLVE Follow the conceptual plan. Begin with mol NaOH and convert to L solution.	**SOLUTION** $0.255\ \cancel{\text{mol NaOH}} \times \dfrac{1\text{ L solution}}{0.125\ \cancel{\text{mol NaOH}}} = 2.04\text{ L solution}$

CHECK The units of the answer (L) are correct. The magnitude is reasonable because the solution contains 0.125 mol per liter. Therefore, roughly 2 L contains the given amount of moles (0.255 mol).

FOR PRACTICE 4.6

How many grams of sucrose ($C_{12}H_{22}O_{11}$) are in 1.55 L of 0.758 M sucrose solution?

FOR MORE PRACTICE 4.6

How many mL of a 0.155 M KCl solution contain 2.55 g KCl?

CONCEPTUAL CONNECTION 4.5 SOLUTIONS

If 25 g of salt dissolves in 251 g of water, what is the mass of the resulting solution?

(a) 251 g **(b)** 276 g **(c)** 226 g

When diluting acids, always add the concentrated acid to the water. Never add water to concentrated acid solutions, as the heat generated may cause the concentrated acid to splatter and burn your skin.

Solution Dilution

To save space in laboratory storerooms, solutions are often stored in concentrated forms called **stock solutions**. For example, hydrochloric acid is often stored as a 12 M stock solution. However, many lab procedures call for hydrochloric acid solutions that are much less concentrated, so we must dilute the stock solution to the required concentration. How do we know how much of the stock solution to use? The easiest way to solve dilution problems is to use the following dilution equation:

$$M_1V_1 = M_2V_2 \qquad [4.1]$$

where M_1 and V_1 are the molarity and volume of the initial concentrated solution, and M_2 and V_2 are the molarity and volume of the final diluted solution. This equation works because the molarity multiplied by the volume gives the number of moles of solute, which is the same in both solutions.

$$M_1V_1 = M_2V_2$$

$$\text{mol}_1 = \text{mol}_2$$

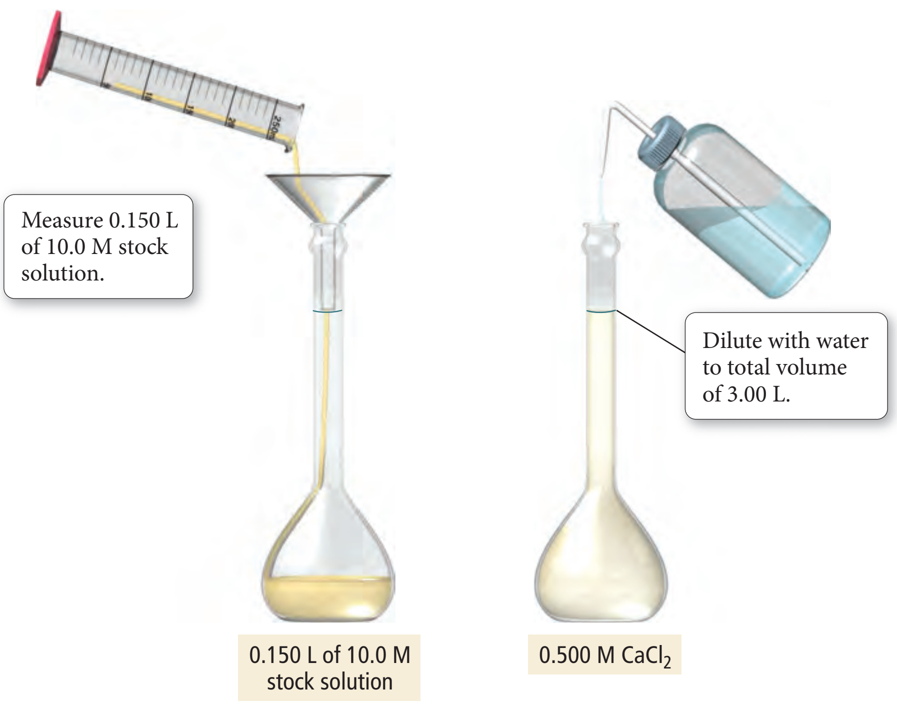
In other words, the number of moles of solute does not change when we dilute a solution.

For example, suppose a laboratory procedure calls for 3.00 L of a 0.500 M $CaCl_2$ solution. How should we prepare this solution from a 10.0 M stock solution? We can solve Equation 4.1 for V_1, the volume of the stock solution required for the dilution, and substitute in the correct values to calculate it.

$$\begin{aligned} M_1V_1 &= M_2V_2 \\ V_1 &= \frac{M_2V_2}{M_1} \\ &= \frac{0.500\ \cancel{\text{mol/L}} \times 3.00\ \text{L}}{10.0\ \cancel{\text{mol/L}}} \\ &= 0.150\ \text{L} \end{aligned}$$

We make the solution by adding enough water to 0.150 L of the stock solution to create a total volume of 3.00 L (V_2). The resulting solution will be 0.500 M in $CaCl_2$ (FIGURE 4.5▼).

Diluting a Solution



$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\ \cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\ \cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

▲ FIGURE 4.5 **Preparing 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution**

EXAMPLE 4.7 Solution Dilution

To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

SORT You are given the initial volume, initial concentration, and final concentration of a solution, and you need to find the final volume.

GIVEN $V_1 = 0.200$ L
$M_1 = 15.0$ M
$M_2 = 3.00$ M

FIND V_2

STRATEGIZE Equation 4.1 relates the initial and final volumes and concentrations for solution dilution problems. You are asked to find V_2. The other quantities (V_1, M_1, and M_2) are all given in the problem.

CONCEPTUAL PLAN

$$V_1, M_1, M_2 \xrightarrow[M_1V_1 = M_2V_2]{} V_2$$

RELATIONSHIPS USED

$M_1V_1 = M_2V_2$

SOLVE Begin with the solution dilution equation and solve it for V_2.
Substitute in the required quantities and calculate V_2.
Make the solution by diluting 0.200 L of the stock solution to a total volume of 1.00 L (V_2). The resulting solution has a concentration of 3.00 M.

SOLUTION $M_1V_1 = M_2V_2$

$$V_2 = \frac{M_1V_1}{M_2}$$

$$= \frac{15.0 \text{ } \cancel{\text{mol/L}} \times 0.200 \text{ L}}{3.00 \text{ } \cancel{\text{mol/L}}}$$

$$= 1.00 \text{ L}$$

CHECK The final units (L) are correct. The magnitude of the answer is reasonable because the solution is diluted from 15.0 M to 3.00 M, a factor of five. Therefore the volume should increase by a factor of five.

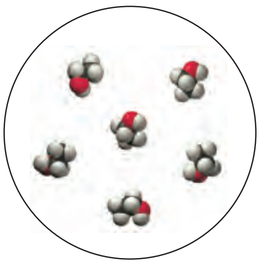
FOR PRACTICE 4.7

To what volume (in mL) should you dilute 100.0 mL of a 5.00 M $CaCl_2$ solution to obtain a 0.750 M $CaCl_2$ solution?

FOR MORE PRACTICE 4.7

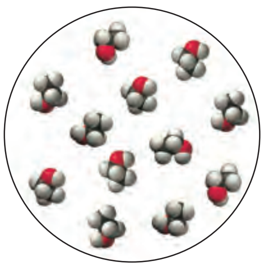
What volume of a 6.00 M $NaNO_3$ solution do you need to make 0.525 L of a 1.20 M $NaNO_3$ solution?

CONCEPTUAL CONNECTION 4.6 SOLUTION DILUTION

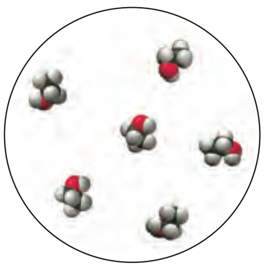


The image at left represents a small volume within 500 mL of an aqueous ethanol (CH_3CH_2OH) solution. (The water molecules have been omitted for clarity.)

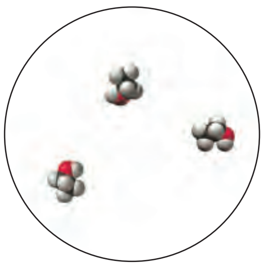
Which of the following images best represents the same volume of the solution after adding an additional 500 mL of water?



(a)



(b)



(c)

Solution Stoichiometry

In Section 4.2 we discussed how we can use the coefficients in chemical equations as conversion factors between the amounts of reactants (in moles) and the amounts of products (in moles). In reactions involving aqueous reactants and products, it is often convenient to specify quantities in terms of volumes and concentrations. We can then use these quantities to calculate the amounts in moles of reactants or products and use the stoichiometric coefficients to convert these to amounts of other reactants or products. The general conceptual plan for these kinds of calculations is:

$$\text{Volume A} \longrightarrow \begin{array}{c}\text{Amount A}\\ \text{(in moles)}\end{array} \longrightarrow \begin{array}{c}\text{Amount B}\\ \text{(in moles)}\end{array} \longrightarrow \text{Volume B}$$

We use molarity to convert between solution volumes and amount of solute, and we use the stoichiometric coefficients from the balanced chemical equation to convert between amounts of A and B. Example 4.8 demonstrates solution stoichiometry.

EXAMPLE 4.8 Solution Stoichiometry

What volume (in L) of 0.150 M KCl solution is required to completely react with 0.150 L of a 0.175 M $Pb(NO_3)_2$ solution according to the following balanced chemical equation?

$$2\,KCl(aq) + Pb(NO_3)_2(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$$

SORT You are given the volume and concentration of a $Pb(NO_3)_2$ solution.
You are asked to find the volume of KCl solution (of a given concentration) required to react with it.

GIVEN 0.150 L of $Pb(NO_3)_2$ solution, 0.175 M $Pb(NO_3)_2$ solution, 0.150 M KCl solution

FIND volume KCl solution (in L)

STRATEGIZE The conceptual plan has the form: volume A ⟶ amount A (in moles) ⟶ amount B (in moles) ⟶ volume B. Use the molar concentrations of the KCl and $Pb(NO_3)_2$ solutions as conversion factors between the number of moles of reactants in these solutions and their volumes. Use the stoichiometric coefficients from the balanced equation to convert between number of moles of $Pb(NO_3)_2$ and number of moles of KCl.

CONCEPTUAL PLAN

$$\text{L } Pb(NO_3)_2 \text{ solution} \xrightarrow{\frac{0.175 \text{ mol } Pb(NO_3)_2}{1 \text{ L } Pb(NO_3)_2 \text{ solution}}} \text{mol } Pb(NO_3)_2 \xrightarrow{\frac{2 \text{ mol KCl}}{1 \text{ mol } Pb(NO_3)_2}}$$

$$\text{mol KCl} \xrightarrow{\frac{1 \text{ L KCl solution}}{0.150 \text{ mol KCl}}} \text{L KCl solution}$$

RELATIONSHIPS USED

$$M\,[Pb(NO_3)_2] = \frac{0.175 \text{ mol } Pb(NO_3)_2}{1 \text{ L } Pb(NO_3)_2 \text{ solution}}$$

$$2 \text{ mol KCl} : 1 \text{ mol } Pb(NO_3)_2$$

$$M\,(KCl) = \frac{0.150 \text{ mol KCl}}{1 \text{ L KCl solution}}$$

SOLVE Begin with L $Pb(NO_3)_2$ solution and follow the conceptual plan to arrive at L KCl solution.

SOLUTION

$$0.150 \cancel{\text{ L } Pb(NO_3)_2 \text{ solution}} \times \frac{0.175 \cancel{\text{ mol } Pb(NO_3)_2}}{1 \cancel{\text{ L } Pb(NO_3)_2 \text{ solution}}} \times \frac{2 \cancel{\text{ mol KCl}}}{1 \cancel{\text{ mol } Pb(NO_3)_2}} \times \frac{1 \text{ L KCl solution}}{0.150 \cancel{\text{ mol KCl}}}$$

$$= 0.350 \text{ L KCl solution}$$

—Continued on the next page

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CHECK The final units (L KCl solution) are correct. The magnitude (0.350 L) is reasonable because the reaction stoichiometry requires 2 mol of KCl per mole of $Pb(NO_3)_2$. Since the concentrations of the two solutions are not very different (0.150 M compared to 0.175 M), the volume of KCl required should be roughly two times the 0.150 L of $Pb(NO_3)_2$ given in the problem.

FOR PRACTICE 4.8

What volume (in mL) of a 0.150 M HNO_3 solution is required to completely react with 35.7 mL of a 0.108 M Na_2CO_3 solution according to the following balanced chemical equation?

$$Na_2CO_3(aq) + 2\ HNO_3(aq) \longrightarrow 2\ NaNO_3(aq) + CO_2(g) + H_2O(l)$$

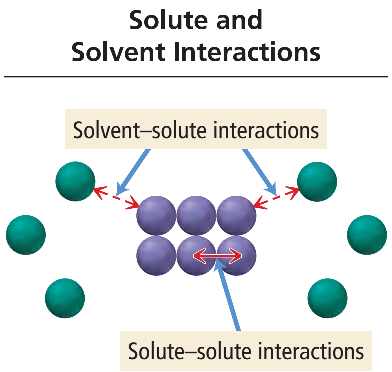
FOR MORE PRACTICE 4.8

In the reaction from For Practice 4.8, what mass (in grams) of carbon dioxide is formed?

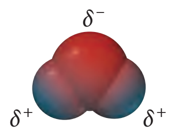
4.5 Types of Aqueous Solutions and Solubility

Consider two familiar aqueous solutions: salt water and sugar water. Salt water is a homogeneous mixture of NaCl and H_2O, and sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O. You may have made these solutions yourself by adding solid table salt or solid sugar to water. As you stir either of these two substances into the water, it seems to disappear. However, you know that the original substance is still present because the mixture tastes salty or sweet. How do solids such as salt and sugar dissolve in water?

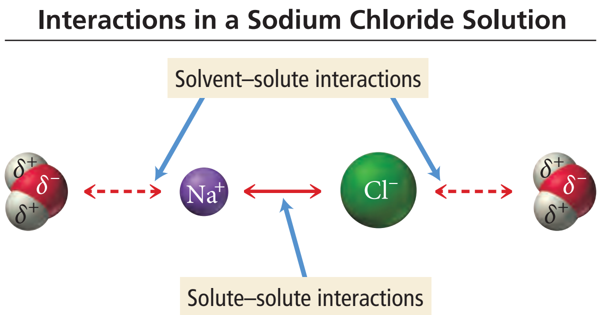
When a solid is put into a liquid solvent, the attractive forces that hold the solid together (the solute–solute interactions) compete with the attractive forces between the solvent molecules and the particles that compose the solid (the solvent–solute interactions), as shown in FIGURE 4.6◄. For example, when sodium chloride is put into water, there is a competition between the attraction of Na^+ cations and Cl^- anions to each other (due to their opposite charges) and the attraction of Na^+ and Cl^- to water molecules. The attraction of Na^+ and Cl^- to water is based on the *polar nature* of the water molecule. For reasons we discuss later in this book (Section 9.6), the oxygen atom in water is electron-rich, giving it a partial negative charge (δ^-), as shown in FIGURE 4.7◄. The hydrogen atoms, in contrast, are electron-poor, giving them a partial positive charge (δ^+). As a result, the positively charged sodium ions are strongly attracted to the oxygen side of the water molecule (which has a partial negative charge), and the negatively charged chloride ions are attracted to the hydrogen side of the water molecule (which has a partial positive charge), as shown in FIGURE 4.8▼. In the case of NaCl, the attraction between the separated ions and the water molecules overcomes the attraction of sodium and chloride ions to each other, and the sodium chloride dissolves in the water (FIGURE 4.9►).



▲ FIGURE 4.6 **Solute and Solvent Interactions** When a solid is put into a solvent, the interactions between solvent and solute particles compete with the interactions among the solute particles themselves.

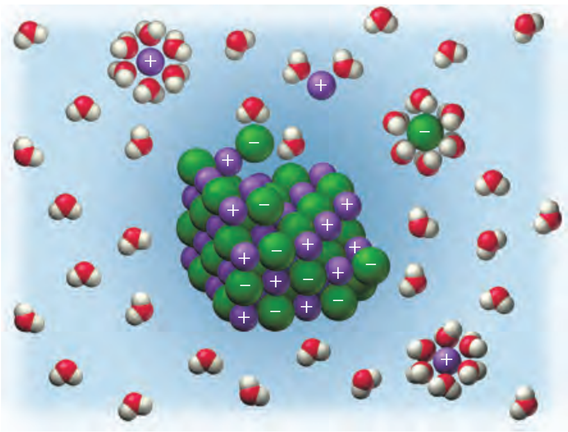


▲ FIGURE 4.7 **Charge Distribution in a Water Molecule** An uneven distribution of electrons within the water molecule causes the oxygen side of the molecule to have a partial negative charge and the hydrogen side to have a partial positive charge.



▲ FIGURE 4.8 **Solute and Solvent Interactions in a Sodium Chloride Solution** When sodium chloride is added to water, the attraction of Na^+ and Cl^- ions to water molecules competes with the attraction among the oppositely charged ions themselves.

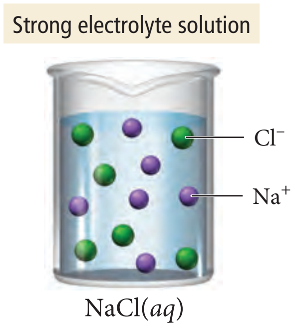
Dissolution of an Ionic Compound



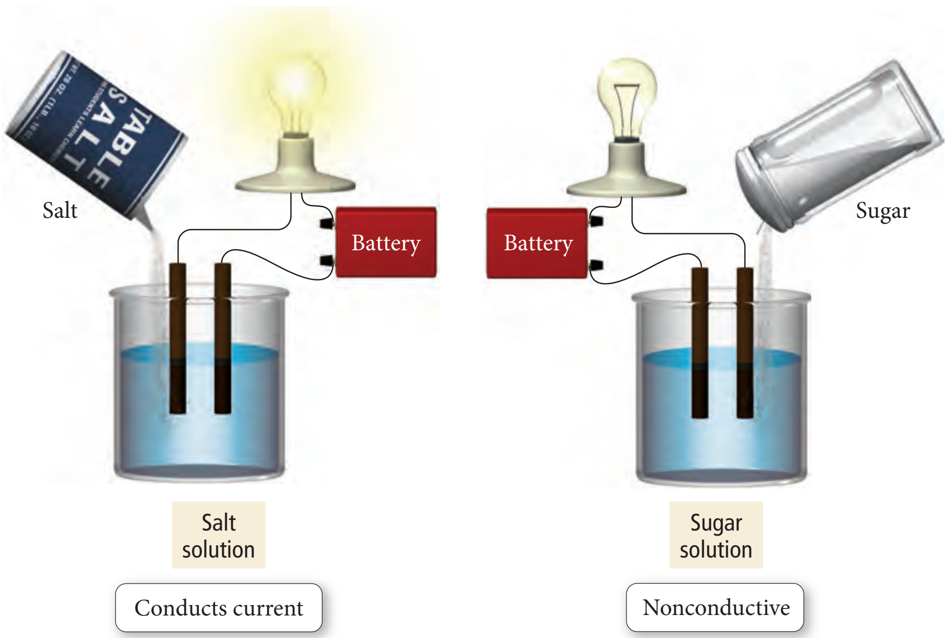
▲ **FIGURE 4.9 Sodium Chloride Dissolving in Water** The attraction between water molecules and the ions of sodium chloride allows NaCl to dissolve in the water.

Electrolyte and Nonelectrolyte Solutions

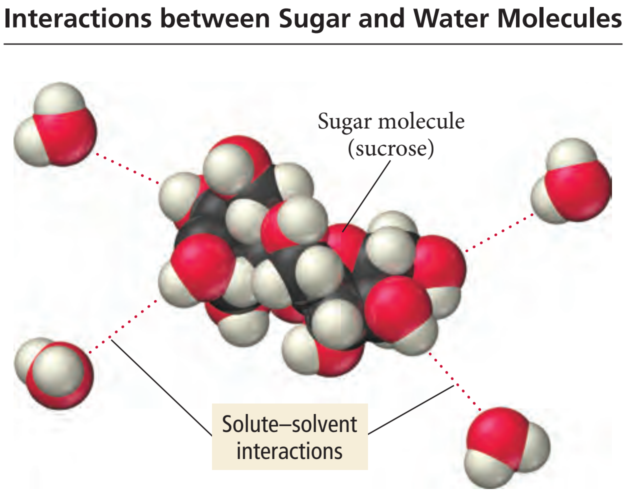
The difference in the way that salt (an ionic compound) and sugar (a molecular compound) dissolve in water illustrates a fundamental difference between types of solutions. As FIGURE 4.10▼ shows, a salt solution conducts electricity, while a sugar solution does not. As we just saw with sodium chloride, ionic compounds dissociate into their component ions when they dissolve in water. An NaCl solution, represented as NaCl(aq), does not contain NaCl units, only Na^+ ions and Cl^- ions. The dissolved ions act as charge carriers, allowing the solution to conduct electricity. Substances that dissolve in water to form solutions that conduct electricity are **electrolytes**. Substances such as sodium chloride that completely dissociate into ions when they dissolve in water are **strong electrolytes**, and the resulting solutions are strong electrolyte solutions.



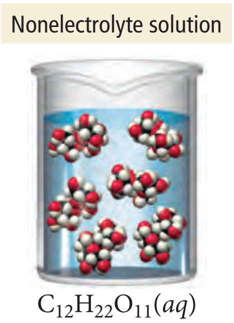
Electrolyte and Nonelectrolyte Solutions



▲ **FIGURE 4.10 Electrolyte and Nonelectrolyte Solutions** A solution of salt (an electrolyte) conducts electrical current. A solution of sugar (a nonelectrolyte) does not.



▶ **FIGURE 4.11 Sugar and Water Interactions** Partial charges on sugar molecules and water molecules (which we will discuss more fully in Chapter 11) result in attractions between the sugar molecules and the water molecules.

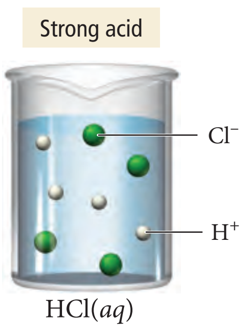


In contrast to sodium chloride, sugar is a molecular compound. Most molecular compounds—with the important exception of acids, which we discuss in the next paragraph—dissolve in water as intact molecules. Sugar dissolves because the attraction between sugar molecules and water molecules—both of which contain a distribution of electrons that results in partial positive and partial negative charges—overcomes the attraction of sugar molecules to each other (FIGURE 4.11▲). However, in contrast to a sodium chloride solution (which is composed of dissociated ions), a sugar solution is composed of intact $C_{12}H_{22}O_{11}$ molecules homogeneously mixed with the water molecules. Compounds such as sugar that do not dissociate into ions when dissolved in water are **nonelectrolytes**, and the resulting solutions—*nonelectrolyte solutions*—do not conduct electricity.

Unlike soluble ionic compounds, which contain ions and therefore *dissociate* in water, acids are molecular compounds that *ionize* in water.

Acids, which we first discussed in Section 3.6, are molecular compounds that ionize—form ions—when they dissolve in water. For example, HCl is a molecular compound that ionizes into H^+ and Cl^- when it dissolves in water. Hydrochloric acid is an example of a **strong acid**, one that completely ionizes in solution. Since they completely ionize in solution, strong acids are also strong electrolytes. We represent the complete ionization of a strong acid with a single reaction arrow between the acid and its ionized form:

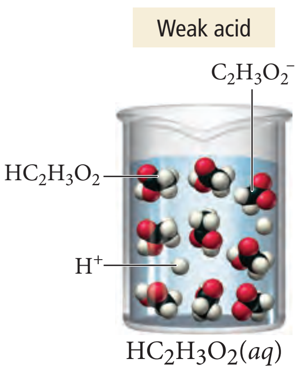
$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$



Many acids are **weak acids**; they do not completely ionize in water. For example, acetic acid ($HC_2H_3O_2$), the acid in vinegar, is a weak acid. A solution of a weak acid is composed mostly of the nonionized acid—only a small percentage of the acid molecules ionize. We represent the partial ionization of a weak acid with opposing half arrows between the reactants and products:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

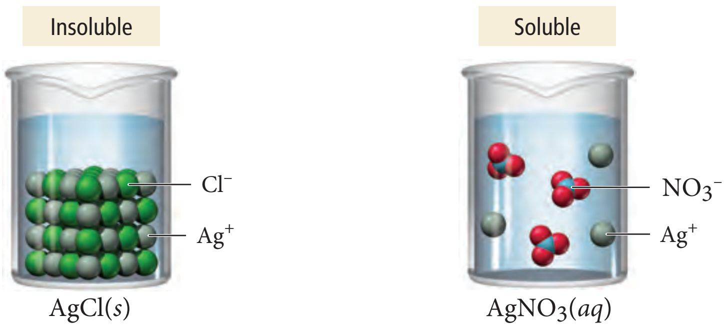
Weak acids are **weak electrolytes**, and the resulting solutions—*weak electrolyte solutions*—conduct electricity only weakly.



The Solubility of Ionic Compounds

As we have just discussed, when an ionic compound dissolves in water, the resulting solution contains—not the intact ionic compound itself—but its component ions dissolved in water. However, not all ionic compounds dissolve in water. If we add AgCl to water, for example, it remains solid and appears as a white powder at the bottom of the beaker.

In general, we say that a compound is **soluble** if it dissolves in water and **insoluble** if it does not. However, these classifications are a bit of an oversimplification. In reality, solubility is a continuum. Compounds exhibit a very wide range of solubilities, and even "insoluble" compounds dissolve to some extent, though usually orders of magnitude less than soluble compounds. For example, silver nitrate is soluble. If we mix solid $AgNO_3$ with water, it dissolves and forms a strong electrolyte solution. Silver chloride, on the other hand, is almost completely insoluble. If we mix solid AgCl with water, virtually all of it remains as a solid within the liquid water.



▲ *AgCl does not significantly dissolve in water; it remains as a white powder at the bottom of the beaker.*

Whether a particular compound is soluble or insoluble depends on several factors. In Section 12.3, we will examine more closely the energy associated with solution formation. For now, we can follow a set of empirical rules that has been inferred from observations of many ionic compounds. Table 4.1 summarizes these *solubility rules*.

For example, the solubility rules state that compounds containing the sodium ion are soluble. That means that compounds such as NaBr, $NaNO_3$, Na_2SO_4, NaOH, and Na_2CO_3 dissolve in water to form strong electrolyte solutions. Similarly, the solubility rules state that compounds containing the NO_3^- ion are soluble. That means that compounds such as $AgNO_3$, $Pb(NO_3)_2$, $NaNO_3$, $Ca(NO_3)_2$, and $Sr(NO_3)_2$ dissolve in water to form strong electrolyte solutions.

Notice that when compounds containing polyatomic ions such as NO_3^- dissolve, the polyatomic ions remain as intact units when they dissolve.

The solubility rules also state that, with some exceptions, compounds containing the CO_3^{2-} ion are insoluble. Therefore, compounds such as $CuCO_3$, $CaCO_3$, $SrCO_3$, and $FeCO_3$ do not dissolve in water. Note that the solubility rules include many exceptions.

TABLE 4.1 Solubility Rules for Ionic Compounds in Water

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li^+, Na^+, K^+, and NH_4^+	None
NO_3^- and $C_2H_3O_2^-$	None
Cl^-, Br^-, and I^-	When these ions pair with Ag^+, Hg_2^{2+}, or Pb^{2+}, the resulting compounds are insoluble.
SO_4^{2-}	When SO_4^{2-} pairs with Sr^{2+}, Ba^{2+}, Pb^{2+}, Ag^+, or Ca^{2+}, the resulting compound is insoluble.
Compounds Containing the Following Ions Are Generally Insoluble	**Exceptions**
OH^- and S^{2-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble. When S^{2-} pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble. When OH^- pairs with Ca^{2+}, Sr^{2+}, or Ba^{2+}, the resulting compound is soluble.
CO_3^{2-} and PO_4^{3-}	When these ions pair with Li^+, Na^+, K^+, or NH_4^+, the resulting compounds are soluble.

EXAMPLE 4.9 Predicting Whether an Ionic Compound Is Soluble

Is each compound soluble or insoluble?

(a) $PbCl_2$ (b) $CuCl_2$ (c) $Ca(NO_3)_2$ (d) $BaSO_4$

SOLUTION

(a) Insoluble. Compounds containing Cl^- are normally soluble, but Pb^{2+} is an exception.

(b) Soluble. Compounds containing Cl^- are normally soluble and Cu^{2+} is not an exception.

(c) Soluble. Compounds containing NO_3^- are always soluble.

(d) Insoluble. Compounds containing SO_4^{2-} are normally soluble, but Ba^{2+} is an exception.

FOR PRACTICE 4.9

Is each compound soluble or insoluble?

(a) NiS (b) $Mg_3(PO_4)_2$ (c) Li_2CO_3 (d) NH_4Cl

KEY CONCEPT VIDEO
Reactions in Solution

4.6 Precipitation Reactions

Have you ever taken a bath in hard water? Hard water contains dissolved ions such as Ca^{2+} and Mg^{2+} that diminish the effectiveness of soap. These ions react with soap to form a gray curd that may appear as a "bathtub ring" after you drain the tub. Hard water is particularly troublesome when washing clothes. Imagine how a white shirt would look covered with the gray curd from the bathtub. Consequently, most laundry detergents include substances designed to remove Ca^{2+} and Mg^{2+} from the laundry mixture. The most common substance used for this purpose is sodium carbonate, which dissolves in water to form sodium cations (Na^+) and carbonate (CO_3^{2-}) anions.

$$Na_2CO_3(aq) \longrightarrow 2\ Na^+(aq) + CO_3^{2-}(aq)$$

◀ *The reaction of ions in hard water with soap produces a gray curd that is visible after bathwater is drained.*

Sodium carbonate is often called washing soda.

Sodium carbonate is soluble, but calcium carbonate and magnesium carbonate are not (see the solubility rules in Table 4.1). Consequently, the carbonate anions react with dissolved Mg^{2+} and Ca^{2+} ions in hard water to form solids that *precipitate* from (come out of) solution:

$$Mg^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow MgCO_3(s)$$

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CaCO_3(s)$$

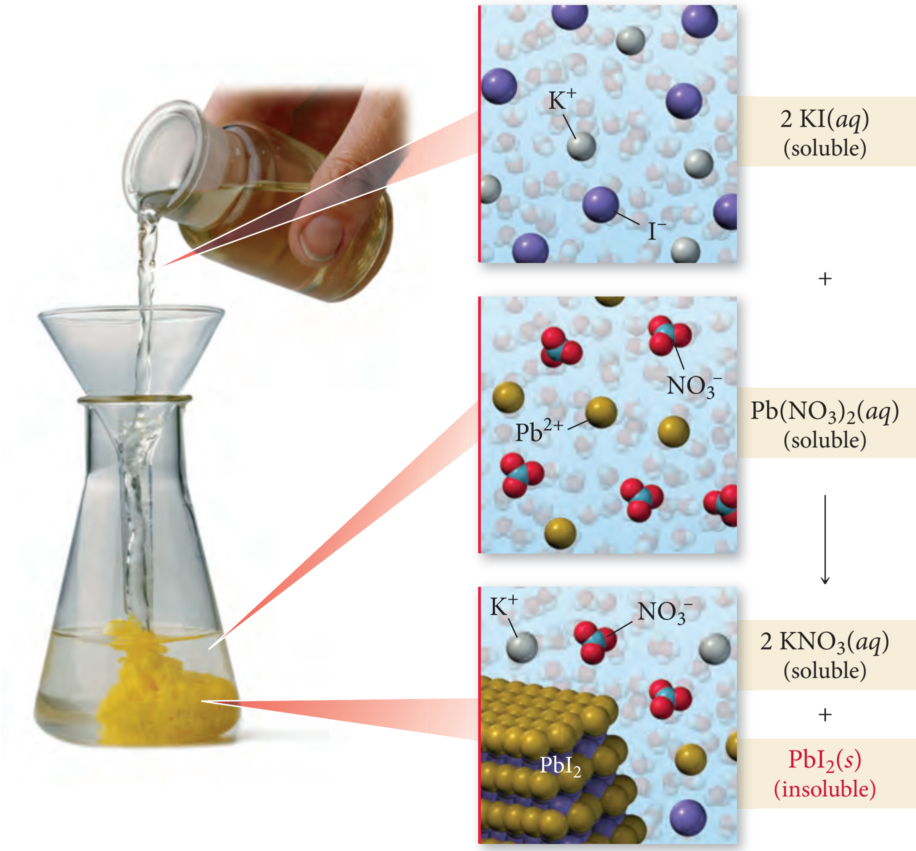
The precipitation of these ions prevents their reaction with the soap, eliminating curd and preventing white shirts from turning gray.

Precipitation Reaction

$$2\,KI(aq) + Pb(NO_3)_2(aq) \longrightarrow 2\,KNO_3(aq) + PbI_2(s)$$

(soluble) (soluble) (soluble) (insoluble)

When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.



◀ **FIGURE 4.12 Precipitation of Lead(II) Iodide** When a potassium iodide solution is mixed with a lead(II) nitrate solution, a yellow lead(II) iodide precipitate forms.

The reactions between CO_3^{2-} and Mg^{2+} and Ca^{2+} are examples of **precipitation reactions**, reactions in which a solid or **precipitate** forms upon mixing two solutions. Precipitation reactions are common in chemistry. As another example, consider potassium iodide and lead(II) nitrate. Each of these compounds forms colorless, strong electrolyte solutions when dissolved in water. When the two solutions are combined, however, a brilliant yellow precipitate forms (FIGURE 4.12▲). We can describe this precipitation reaction with the chemical equation:

$$2\,KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2\,KNO_3(aq) + PbI_2(s)$$

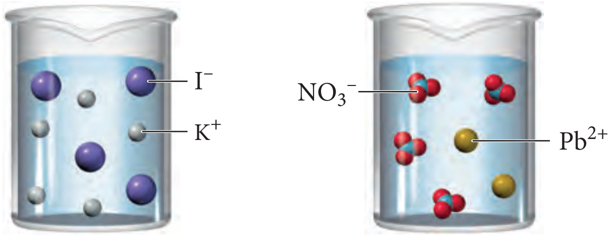
Precipitation reactions do not always occur when two aqueous solutions are mixed. For example, if we combine solutions of KI(*aq*) and NaCl(*aq*), nothing happens.

$$KI(aq) + NaCl(aq) \rightarrow \text{NO REACTION}$$

The key to predicting precipitation reactions is to understand that only *insoluble* compounds form precipitates. In a precipitation reaction, two solutions containing soluble compounds combine and an insoluble compound precipitates. For example, consider the precipitation reaction described previously:

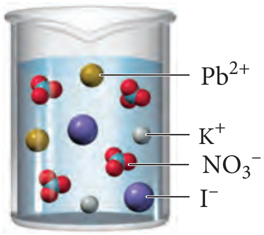
$$\underset{\text{soluble}}{2\,KI(aq)} + \underset{\text{soluble}}{Pb(NO_3)_2(aq)} \rightarrow \underset{\text{insoluble}}{PbI_2(s)} + \underset{\text{soluble}}{2\,KNO_3(aq)}$$

KI and $Pb(NO_3)_2$ are both soluble, but the precipitate, PbI_2, is insoluble. Before mixing, KI(*aq*) and $Pb(NO_3)_2$(*aq*) are both dissociated in their respective solutions.



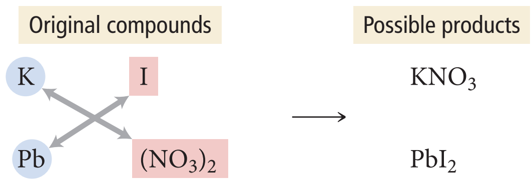
KI(*aq*) $Pb(NO_3)_2$(*aq*)

The instant that the solutions are mixed, all four ions are present.

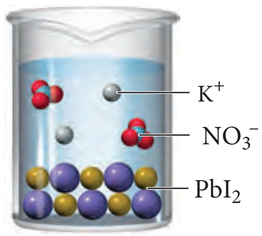


KI(*aq*) and $Pb(NO_3)_2$(*aq*)

Two new compounds—one or both of which might be insoluble—are now possible. The cation from each compound can pair with the anion from the other compound to form possibly insoluble products (we will learn more about why this happens in Chapter 12).



If the possible products are both soluble, then no reaction occurs. If one or both of the possible products are insoluble, a precipitation reaction occurs. In this case, KNO_3 is soluble, but PbI_2 is insoluble. Consequently, PbI_2 precipitates as shown here:



PbI_2(*s*) and KNO_3(*aq*)

To predict whether a precipitation reaction will occur when two solutions are mixed and to write an equation for the reaction, use the procedure that follows. The steps are outlined in the left column, and examples of applying the procedure are shown in the center and right columns in Examples 4.10 and 4.11.

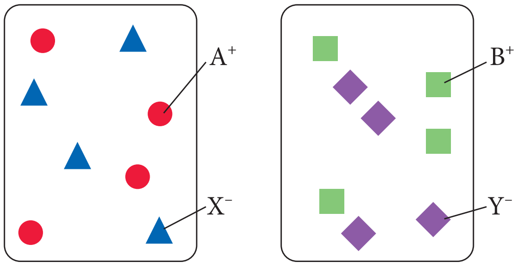
PROCEDURE FOR...	EXAMPLE 4.10	EXAMPLE 4.11
Writing Equations for Precipitation Reactions	**Writing Equations for Precipitation Reactions**	**Writing Equations for Precipitation Reactions**
	Write an equation for the precipitation reaction that occurs (if any) when solutions of potassium carbonate and nickel(II) chloride are mixed.	Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium nitrate and lithium sulfate are mixed.
1. Write the formulas of the two compounds being mixed as reactants in a chemical equation.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$	$NaNO_3(aq) + Li_2SO_4(aq) \longrightarrow$
2. Below the equation, write the formulas of the products that could form from the reactants. Obtain these by combining the cation from each reactant with the anion from the other. Make sure to write correct formulas for these ionic compounds, as described in Section 3.5.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$ Possible products KCl $NiCO_3$	$NaNO_3(aq) + Li_2SO_4(aq) \longrightarrow$ Possible products $LiNO_3$ Na_2SO_4
3. Consult the solubility rules to determine whether any of the possible products are insoluble.	KCl is soluble. (Compounds containing Cl^- are usually soluble, and K^+ is not an exception.) $NiCO_3$ is insoluble. (Compounds containing CO_3^{2-} are usually insoluble, and Ni^{2+} is not an exception.)	$LiNO_3$ is soluble. (Compounds containing NO_3^- are always soluble.) Na_2SO_4 is soluble. (Compounds containing SO_4^{2-} are usually soluble, and Na^+ is not an exception.)
4. If all of the possible products are soluble, there will be no precipitate. Write NO REACTION after the arrow.	Since this example has an insoluble product, we proceed to the next step.	Since this example has no insoluble product, there is no reaction. $NaNO_3(aq) + Li_2SO_4(aq) \longrightarrow$ NO REACTION
5. If any of the possible products are insoluble, write their formulas as the products of the reaction using (*s*) to indicate solids. Write any soluble products with (*aq*) to indicate aqueous products.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$ $NiCO_3(s) + KCl(aq)$	
6. Balance the equation. Remember to adjust only coefficients, not subscripts.	$K_2CO_3(aq) + NiCl_2(aq) \longrightarrow$ $NiCO_3(s) + 2\ KCl(aq)$	
	FOR PRACTICE 4.10 Write an equation for the precipitation reaction that occurs (if any) when solutions of ammonium chloride and iron(III) nitrate are mixed.	**FOR PRACTICE 4.11** Write an equation for the precipitation reaction that occurs (if any) when solutions of sodium hydroxide and copper(II) bromide are mixed.

CONCEPTUAL CONNECTION 4.7 PRECIPITATION REACTIONS

Consider the generic ionic compounds with the formulas AX and BY and the following solubility rules:

AX soluble; BY soluble; AY soluble; BX insoluble

Let circles represent A^+ ions, squares represent B^+ ions, triangles represent X^- ions, and diamonds represent Y^- ions. We can then represent solutions of the two compounds (AX and BY) as follows:



AX(aq) BY(aq)

Draw a molecular-level representation showing the result of mixing the two solutions and write an equation to represent the reaction.

4.7 Representing Aqueous Reactions: Molecular, Ionic, and Complete Ionic Equations

Consider the following equation for a precipitation reaction:

$$Pb(NO_3)_2(aq) + 2\,KCl(aq) \longrightarrow PbCl_2(s) + 2\,KNO_3(aq)$$

This equation is a **molecular equation**, an equation that shows the complete neutral formulas for each compound in the reaction as if they existed as molecules. However, in actual solutions of soluble ionic compounds, dissociated substances are present as ions. Equations for reactions occurring in aqueous solution can be written to better show the dissociated nature of dissolved ionic compounds. For example, we can rewrite this molecular equation as:

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Equations such as this, which list individually all of the ions present as either reactants or products in a chemical reaction, are **complete ionic equations**.

Notice that in the complete ionic equation, some of the ions in solution appear unchanged on both sides of the equation. These ions are called **spectator ions** because they do not participate in the reaction.

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Spectator ions

To simplify the equation, and to show more clearly what is happening, we can omit spectator ions.

$$Pb^{2+}(aq) + 2\,Cl^-(aq) \longrightarrow PbCl_2(s)$$

Equations such as this one, which show only the species that actually change during the reaction, are **net ionic equations**.

As another example, consider the reaction between $HCl(aq)$ and $KOH(aq)$.

$$HCl(aq) + KOH(aq) \longrightarrow H_2O(l) + KCl(aq)$$

Since HCl, KOH, and KCl all exist in solution primarily as independent ions, the complete ionic equation is:

$$H^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \longrightarrow$$
$$H_2O(l) + K^+(aq) + Cl^-(aq)$$

To write the net ionic equation, we remove the spectator ions, those that are unchanged on both sides of the equation.

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{K^+}(aq) + OH^-(aq) \longrightarrow H_2O(l) + \cancel{K^+}(aq) + \cancel{Cl^-}(aq)$$

Spectator ions

The net ionic equation is $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$.

Summarizing: Aqueous Equations

- **A molecular equation** is a chemical equation showing the complete, neutral formulas for every compound in a reaction.
- **A complete ionic equation** is a chemical equation showing all of the species as they are actually present in solution.
- **A net ionic equation** is an equation showing only the species that actually change during the reaction.

EXAMPLE 4.12 Writing Complete Ionic and Net Ionic Equations

Consider the following precipitation reaction occurring in aqueous solution:

$$3\ SrCl_2(aq) + 2\ Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\ LiCl(aq)$$

Write the complete ionic equation and net ionic equation for this reaction.

SOLUTION Write the complete ionic equation by separating aqueous ionic compounds into their constituent ions. The $Sr_3(PO_4)_2(s)$, which precipitates as a solid, remains as one unit.	**Complete ionic equation:** $3\ Sr^{2+}(aq) + 6\ Cl^-(aq) + 6\ Li^+(aq) + 2\ PO_4^{3-}(aq) \longrightarrow$ $Sr_3(PO_4)_2(s) + 6\ Li^+(aq) + 6\ Cl^-(aq)$
Write the net ionic equation by eliminating the spectator ions, those that do not change from one side of the reaction to the other.	**Net ionic equation:** $3\ Sr^{2+}(aq) + 2\ PO_4^{3-}(aq) \longrightarrow Sr_3(PO_4)_2(s)$

FOR PRACTICE 4.12

Consider the following reaction occurring in aqueous solution:

$$2\ HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\ H_2O(l) + BaI_2(aq)$$

Write the complete ionic equation and net ionic equation for this reaction.

FOR MORE PRACTICE 4.12

Write complete ionic and net ionic equations for the following reaction occurring in aqueous solution:

$$2\ AgNO_3(aq) + MgCl_2(aq) \longrightarrow 2\ AgCl(s) + Mg(NO_3)_2(aq)$$

4.8 Acid–Base and Gas-Evolution Reactions

Two other important classes of reactions that occur in aqueous solution are acid–base reactions and gas-evolution reactions. In an **acid–base reaction** (also called a **neutralization reaction**), an acid reacts with a base and the two neutralize each other, producing water (or in some cases a weak electrolyte). In a **gas-evolution reaction**, a gas forms, resulting in bubbling. In both cases, as in precipitation reactions, the reactions occur when the anion from one reactant combines with the cation of the other. In addition, many gas-evolution reactions are also acid–base reactions.

▲ *In a gas-evolution reaction, such as the reaction of hydrochloric acid (HCl) with limestone ($CaCO_3$) to produce carbon dioxide (CO_2), bubbling typically occurs as the gas is released.*

Acid–Base Reactions

Our stomachs contain hydrochloric acid, which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing acid stomach or heartburn. Antacids are over-the-counter medicines that work by reacting with, and neutralizing, stomach acid. Antacids employ different *bases*—substances that produce hydroxide (OH^-) ions in water—as neutralizing agents. Milk of magnesia, for example, contains $Mg(OH)_2$ and Mylanta contains $Al(OH)_3$. All antacids, no matter what base they contain, have the same effect of neutralizing stomach acid through *acid–base reactions* and relieving heartburn.

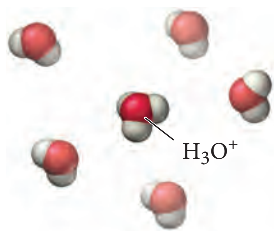
Recall from Chapter 3 that an acid forms H^+ ions in solution, and we just saw that a base is a substance that produces OH^- ions in solution:

- **Acid** **Substance that produces H^+ ions in aqueous solution**
- **Base** **Substance that produces OH^- ions in aqueous solution**

These definitions of acids and bases are called the **Arrhenius definitions**, after Swedish chemist Svante Arrhenius (1859–1927). In Chapter 15, we will learn more general definitions of acid–base behavior, but these are sufficient to describe neutralization reactions.

According to the Arrhenius definition, HCl is an acid because it produces H^+ ions in solution:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$



▲ **FIGURE 4.13 The Hydronium Ion** Protons normally associate with water molecules in solution to form H_3O^+ ions, which in turn interact with other water molecules.

An H^+ ion is a bare proton. Protons associate with water molecules in solution to form **hydronium ions** (FIGURE 4.13◄):

$$H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$$

Chemists use $H^+(aq)$ and $H_3O^+(aq)$ interchangeably to mean the same thing—an H^+ ion dissolved in water. The ionization of HCl and other acids is often written to show the association of the proton with a water molecule to form the hydronium ion.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

As we discussed in Section 4.6, some acids are weak acids—they do not completely ionize in solution. We represent the partial ionization of a weak acid with opposing half arrows.

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

▲ *Acids are found in lemons, limes, and vinegar. Vitamin C and aspirin are also acids.*

Some acids—called **polyprotic acids**—contain more than one ionizable proton and release them sequentially. For example, sulfuric acid, H_2SO_4, is a **diprotic acid**. It is strong in its first ionizable proton, but weak in its second.

$$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$$
$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

According to the Arrhenius definition, NaOH is a base because it produces OH^- ions in solution.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

In analogy to diprotic acids, some bases, such as $Sr(OH)_2$, for example, produce two moles of OH^- per mole of the base.

$$Sr(OH)_2(aq) \rightarrow Sr^{2+}(aq) + 2\ OH^-(aq)$$

Table 4.2 lists common acids and bases. Acids and bases are in many everyday substances.

TABLE 4.2 Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Formic acid	$HCHO_2$ (weak acid)		
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

▲ *Many common household products contain bases.*

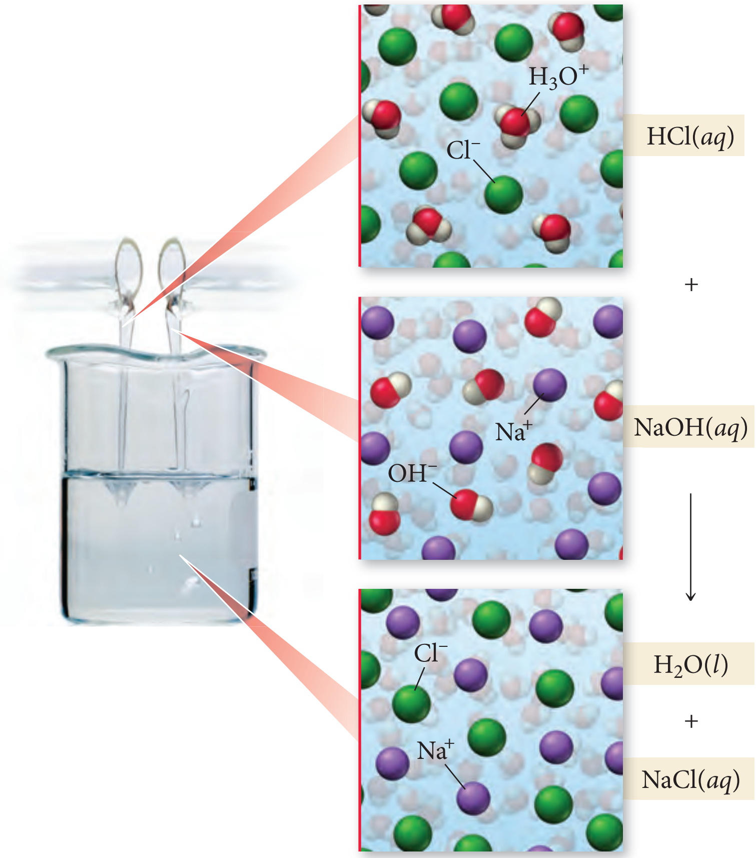
When an acid and a base are mixed, the $H^+(aq)$ from the acid—whether it is weak or strong—combines with the $OH^-(aq)$ from the base to form $H_2O(l)$ (FIGURE 4.14▼). For example, consider the reaction between hydrochloric acid and sodium hydroxide:

$$\underbrace{HCl(aq)}_{\text{Acid}} + \underbrace{NaOH(aq)}_{\text{Base}} \longrightarrow \underbrace{H_2O(l)}_{\text{Water}} + \underbrace{NaCl(aq)}_{\text{Salt}}$$

Acid–Base Reaction

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.



◄ **FIGURE 4.14 Acid–Base Reaction** The reaction between hydrochloric acid and sodium hydroxide forms water and a salt, sodium chloride, which remains dissolved in the solution.

The word *salt* in this sense applies to any ionic compound and is more general than the common usage, which refers only to table salt (NaCl).

Acid–base reactions generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution. The net ionic equation for acid–base reactions involving a strong acid is

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

However, if the acid is a weak acid, the net ionic equation is slightly different. For example, consider the acid–base equation between hydrofluoric acid and sodium hydroxide:

$$\underset{\text{acid}}{HF(aq)} + \underset{\text{base}}{NaOH(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{NaF(aq)}$$

The complete ionic equation and net ionic equation for this reaction are:

$$HF(aq) + Na^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + Na^+(aq) + F^-(aq) \qquad \text{(Complete ionic eqution)}$$

$$HF(aq) + OH^-(aq) \longrightarrow H_2O(l) + F^-(aq) \qquad \text{(Net ionic equation)}$$

Notice that, since HF is a weak acid, it is not shown as being ionized in the ionic equations.

Another example of an acid–base reaction is the reaction between sulfuric acid and potassium hydroxide:

$$\underset{\text{acid}}{H_2SO_4(aq)} + \underset{\text{base}}{2\,KOH(aq)} \longrightarrow \underset{\text{water}}{2\,H_2O(l)} + \underset{\text{salt}}{K_2SO_4(aq)}$$

Again, notice the pattern of acid and base reacting to form water and a salt.

$$\textbf{Acid} + \textbf{Base} \longrightarrow \textbf{Water} + \textbf{Salt} \qquad \text{(acid–base reactions)}$$

When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds from Section 3.5.

EXAMPLE 4.13 Writing Equations for Acid–Base Reactions Involving a Strong Acid

Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous $Ba(OH)_2$.

SOLUTION You must first recognize these substances as an acid and a base. Begin by writing the unbalanced equation in which the acid and the base combine to form water and a salt.	$\underset{\text{acid}}{HI(aq)} + \underset{\text{base}}{Ba(OH)_2(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{BaI_2(aq)}$
Next, balance the equation; this is the molecular equation.	**Molecular equation** $2\,HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + BaI_2(aq)$
Write the net ionic equation by removing the spectator ions.	**Net ionic equation** $2\,H^+(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l)$ or simply $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

FOR PRACTICE 4.13

Write a molecular and a net ionic equation for the reaction that occurs between aqueous HBr and aqueous LiOH.

EXAMPLE 4.14 Writing Equations for Acid–Base Reactions Involving a Weak Acid

Write a molecular equation, ionic equation, and net ionic equation for the reaction between aqueous acetic acid ($HC_2H_3O_2$) and aqueous potassium hydroxide (KOH).

SOLUTION Begin by writing the molecular equation in which the acid and the base combine to form water and a salt. (The equation is already balanced.)	**Molecular equation** $HC_2H_3O_2(aq) + KOH(aq) \longrightarrow H_2O(l) + KC_2H_3O_2(aq)$
Write the complete ionic equation by separating aqueous ionic compounds into their constituent ions. Do not separate $HC_2H_3O_2(aq)$ because it is a weak acid (and a weak electrolyte).	**Complete ionic equation** $HC_2H_3O_2(aq) + K^+(aq) + OH^-(aq) \longrightarrow$ $H_2O(l) + K^+(aq) + C_2H_3O_2^-(aq)$
Write the net ionic equation by eliminating the spectator ions.	**Net ionic equation** $HC_2H_3O_2(aq) + OH^-(aq) \longrightarrow H_2O(l) + C_2H_3O_2^-(aq)$

FOR PRACTICE 4.14

Write the net ionic equation for the reaction between $HCHO_2$ (a weak acid) and NaOH.

Gas-Evolution Reactions

Aqueous reactions that form a gas when two solutions are mixed are *gas-evolution reactions*. Some gas-evolution reactions form a gaseous product directly when the cation of one reactant combines with the anion of the other. For example, when sulfuric acid reacts with lithium sulfide, dihydrogen sulfide gas forms:

$$H_2SO_4(aq) + Li_2S(aq) \longrightarrow \underbrace{H_2S(g)}_{\text{gas}} + H_2SO_4(aq)$$

Other gas-evolution reactions form an intermediate product that then decomposes to form a gas. For example, when aqueous hydrochloric acid is mixed with aqueous sodium bicarbonate, the following reaction occurs (FIGURE 4.15▸):

$$HCl(aq) + NaHCO_3(aq) \longrightarrow H_2CO_3(aq) + NaCl(aq) \longrightarrow H_2O(l) + \underbrace{CO_2(g)}_{\text{gas}} + NaCl(aq)$$

Many gas-evolution reactions such as this one are also acid–base reactions. In Chapter 15 we will learn how ions such as CO_3^{2-} act as bases in aqueous solution.

The intermediate product, H_2CO_3, is not stable and decomposes into H_2O and gaseous CO_2. Other important gas-evolution reactions form H_2SO_3 or NH_4OH as intermediate products:

$$HCl(aq) + NaHSO_3(aq) \longrightarrow H_2SO_3(aq) + NaCl(aq) \longrightarrow H_2O(l) + SO_2(g) + NaCl(aq)$$

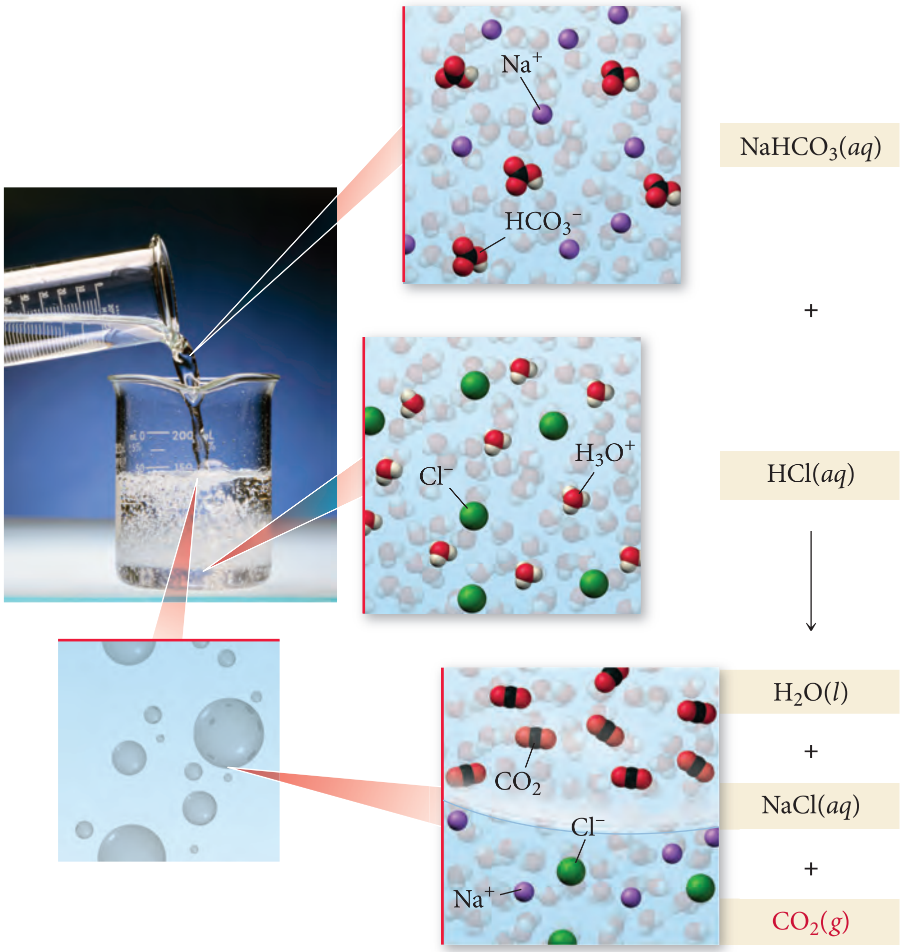
$$NH_4Cl(aq) + NaOH(aq) \longrightarrow NH_4OH(aq) + NaCl(aq) \longrightarrow H_2O(l) + NH_3(g) + NaCl(aq)$$

The intermediate product NH_4OH provides a convenient way to think about this reaction, but the extent to which it actually forms is debatable.

Gas-Evolution Reaction

$$NaHCO_3(aq) + HCl(aq) \longrightarrow H_2O(l) + NaCl(aq) + CO_2(g)$$

When aqueous sodium bicarbonate is mixed with aqueous hydrochloric acid, gaseous CO_2 bubbles are the result of the reaction.



▲ **FIGURE 4.15 Gas-Evolution Reaction** When aqueous hydrochloric acid is mixed with aqueous sodium bicarbonate, gaseous CO_2 bubbles out of the reaction mixture.

Table 4.3 lists the main types of compounds that form gases in aqueous reactions, as well as the gases that form or evolve.

TABLE 4.3 Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2\ HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\ KCl(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2\ HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\ KCl(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2\ HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\ KCl(aq)$
Ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

EXAMPLE 4.15 Writing Equations for Gas-Evolution Reactions

Write a molecular equation for the gas-evolution reaction that occurs when aqueous nitric acid and aqueous sodium carbonate are mixed.

Begin by writing an unbalanced equation in which the cation of each reactant combines with the anion of the other.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2CO_3(aq) + NaNO_3(aq)$
You must recognize that $H_2CO_3(aq)$ decomposes into $H_2O(l)$ and $CO_2(g)$ and write these products into the equation.	$HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$
Finally, balance the equation.	$2\ HNO_3(aq) + Na_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\ NaNO_3(aq)$

FOR PRACTICE 4.15

Write a molecular equation for the gas-evolution reaction that occurs when aqueous hydrobromic acid and aqueous potassium sulfite are mixed.

FOR MORE PRACTICE 4.15

Write a net ionic equation for the reaction that occurs when hydroiodic acid and calcium sulfide are mixed.

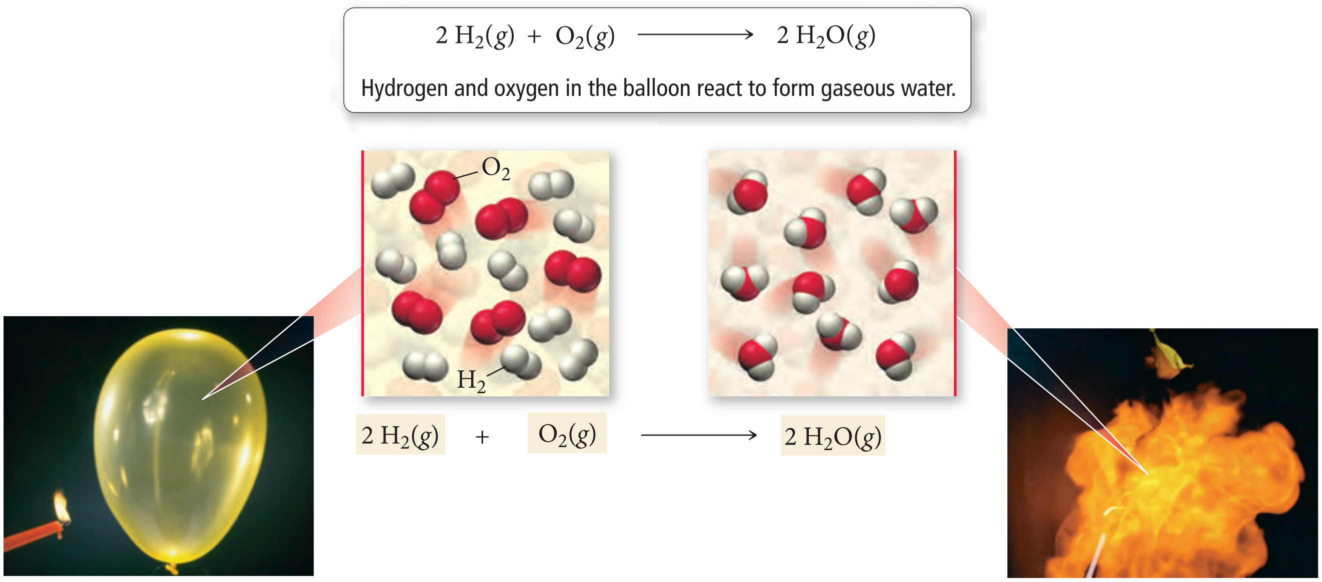
4.9 Oxidation–Reduction Reactions

Oxidation–reduction reactions or **redox reactions** are reactions in which electrons are transferred from one reactant to the other. The rusting of iron, the bleaching of hair, and the production of electricity in batteries involve redox reactions. Many redox reactions involve the reaction of a substance with oxygen (FIGURE 4.16▼):

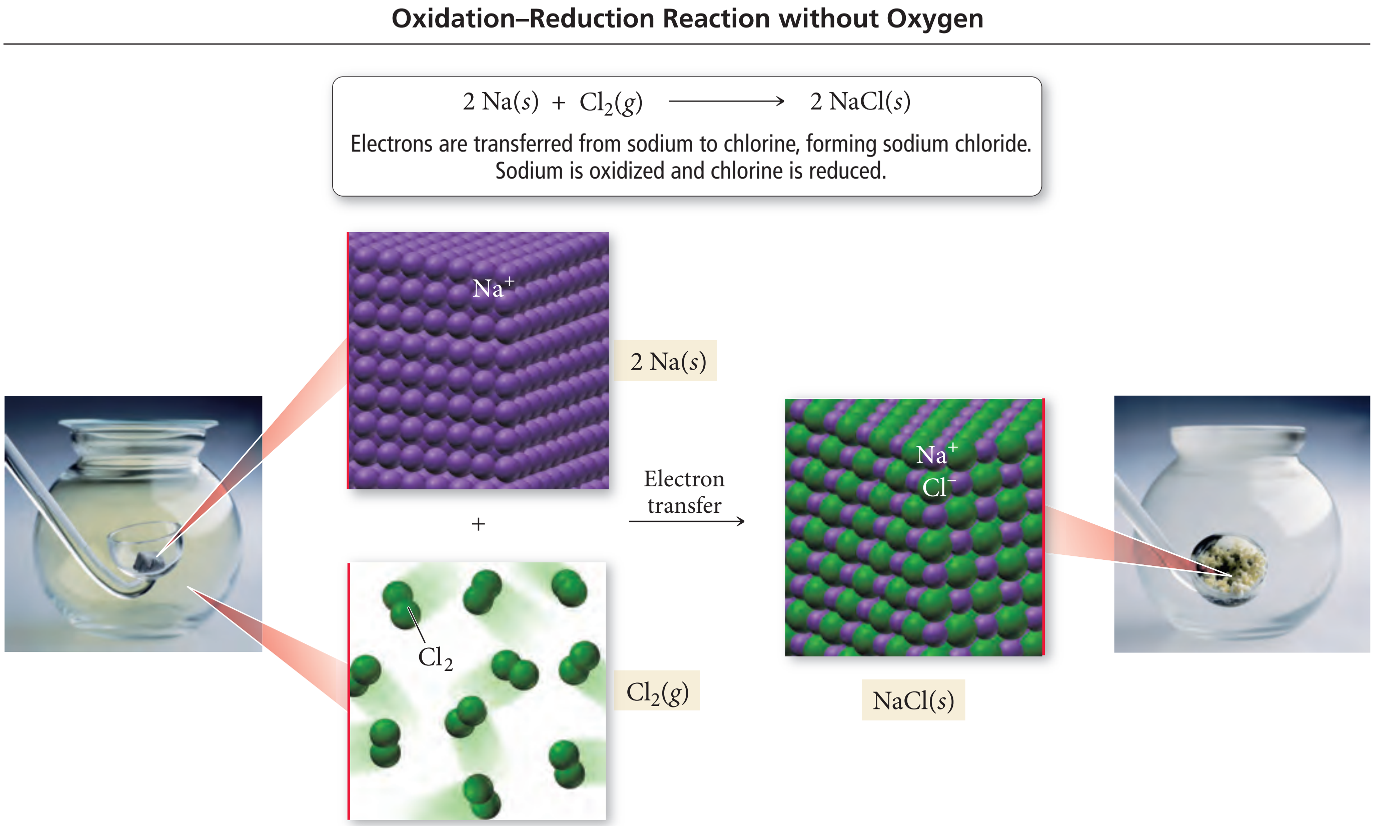
Oxidation–reduction reactions are covered in more detail in Chapter 18.

$$4\ Fe(s) + 3\ O_2(g) \longrightarrow 2\ Fe_2O_3(s) \quad \text{(rusting of iron)}$$
$$2\ C_8H_{18}(l) + 25\ O_2(g) \longrightarrow 16\ CO_2(g) + 18\ H_2O(g) \quad \text{(combustion of octane)}$$
$$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g) \quad \text{(combustion of hydrogen)}$$

Oxidation–Reduction Reaction



▲ FIGURE 4.16 **Oxidation–Reduction Reaction** When heat is applied, the hydrogen in the balloon reacts explosively with oxygen to form gaseous water.



▲ **FIGURE 4.17 Oxidation–Reduction without Oxygen** When sodium reacts with chlorine, electrons are transferred from the sodium to the chlorine, resulting in the formation of sodium chloride. In this redox reaction, sodium is oxidized and chlorine is reduced.

The reaction between sodium and oxygen forms other oxides as well.

A helpful mnemonic is O I L R I G—**O**xidation **I**s **L**oss; **R**eduction **I**s **G**ain.

However, redox reactions need not involve oxygen. Consider, for example, the reaction between sodium and chlorine to form sodium chloride (NaCl), depicted in FIGURE 4.17▲.

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

This reaction is similar to the reaction between sodium and oxygen to form sodium oxide.

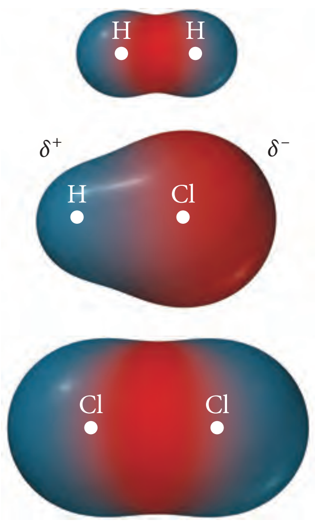
$$4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$$

In both cases, a metal (which has a tendency to lose electrons) reacts with a nonmetal (which has a tendency to gain electrons). In both cases, metal atoms lose electrons to nonmetal atoms. A fundamental definition of **oxidation** is the loss of electrons, and a fundamental definition of **reduction** is the gain of electrons.

The transfer of electrons, however, need not be a *complete* transfer (as occurs in the formation of an ionic compound) for the reaction to qualify as oxidation–reduction. For example, consider the reaction between hydrogen gas and chlorine gas:

$$H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)$$

Even though hydrogen monochloride is a molecular compound with a covalent bond, and even though the hydrogen has not completely transferred its electron to chlorine during the reaction, you can see from the electron density diagrams (FIGURE 4.18◀) that hydrogen has lost some of its electron density—it has *partially* transferred its electron to chlorine. Therefore, in this reaction, hydrogen is oxidized and chlorine is reduced and the reaction is a redox reaction.



▲ **FIGURE 4.18 Redox with Partial Electron Transfer** When hydrogen bonds to chlorine, the electrons are unevenly shared, resulting in an increase of electron density (reduction) for chlorine and a decrease in electron density (oxidation) for hydrogen.

Oxidation States

Identifying a reaction between a metal and a nonmetal as a redox reaction is fairly straightforward because the metal becomes a cation and the nonmetal becomes an anion (electron transfer is obvious). However, how do we identify redox reactions that occur between nonmetals? Chemists have devised a scheme to track electrons before and after a chemical reaction. In this scheme—which is like bookkeeping for electrons—all shared electrons are assigned to the atom that attracts the electrons most strongly. Then a number, called the **oxidation state** or **oxidation number**, is given to each atom based on the electron assignments. In other words, the oxidation number of an atom in a compound is the "charge" it would have if all shared electrons were assigned to the atom with a greater attraction for those electrons.

The ability of an element to attract electrons in a chemical bond is called electronegativity. We will cover electronegativity in more detail in Section 9.6.

For example, consider HCl. Since chlorine attracts electrons more strongly than hydrogen, we assign the two shared electrons in the bond to chlorine; then H (which has lost an electron in our assignment) has an oxidation state of +1, and Cl (which has gained one electron in our assignment) has an oxidation state of −1. Notice that, in contrast to ionic charges, which are usually written with the sign of the charge after the magnitude (1+ and 1−, for example), oxidation states are written with the sign of the charge before the magnitude (+1 and −1, for example). Use the following rules to assign oxidation states to atoms in elements and compounds.

Do not confuse oxidation state with ionic charge. Unlike ionic charge—which is a real property of an ion—the oxidation state of an atom is merely a theoretical (but useful) construct.

Rules for Assigning Oxidation States	Examples	
(These rules are hierarchical. If any two rules conflict, follow the rule that is higher on the list.)		
1. The oxidation state of an atom in a free element is 0.	Cu 0 ox state	Cl_2 0 ox state
2. The oxidation state of a monoatomic ion is equal to its charge.	Ca^{2+} +2 ox state	Cl^- −1 ox state
3. The sum of the oxidation states of all atoms in:		
• A neutral molecule or formula unit is 0.	H_2O 2(H ox state) + 1(O ox state) = 0	
• An ion is equal to the charge of the ion.	NO_3^- 1(N ox state) + 3(O ox state) = −1	
4. In their compounds, metals have positive oxidation states.		
• Group 1A metals *always* have an oxidation state of +1.	NaCl +1 ox state	
• Group 2A metals *always* have an oxidation state of +2.	CaF_2 +2 ox state	
5. In their compounds, nonmetals are assigned oxidation states according to the table that lists oxidation states of nonmetals in the margin. Entries at the top of the table take precedence over entries at the bottom of the table.		

Oxidation States of Nonmetals

Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 −1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 −2 ox state
Group 7A	−1	CCl_4 −1 ox state
Group 6A	−2	H_2S −2 ox state
Group 5A	−3	NH_3 −3 ox state

When assigning oxidation states, keep these points in mind:

- The oxidation state of any given element generally depends on what other elements are present in the compound. (The exceptions are the group 1A and 2A metals, which are *always* +1 and +2, respectively.)
- Rule 3 must always be followed. Therefore, when following the hierarchy shown in rule 5, give priority to the element(s) highest on the list and then assign the oxidation state of the element lowest on the list using rule 3.
- When assigning oxidation states to elements that are not covered by rules 4 and 5 (such as carbon) use rule 3 to deduce their oxidation state once all other oxidation states have been assigned.

EXAMPLE 4.16 Assigning Oxidation States

Assign an oxidation state to each atom in each compound.

(a) Cl_2 (b) Na^+ (c) KF (d) CO_2 (e) SO_4^{2-} (f) K_2O_2

SOLUTION

Since Cl_2 is a free element, the oxidation state of both Cl atoms is 0 (rule 1).	(a) Cl_2 $\underset{0}{Cl}\ \underset{0}{Cl}$
Since Na^+ is a monoatomic ion, the oxidation state of the Na^+ ion is +1 (rule 2).	(b) Na^+ $\underset{+1}{Na^+}$
The oxidation state of K is +1 (rule 4). The oxidation state of F is −1 (rule 5). Since this is a neutral compound, the sum of the oxidation states is 0.	(c) KF $\underset{+1}{K}\underset{-1}{F}$ sum: $+1-1=0$
The oxidation state of oxygen is −2 (rule 5). Deduce the oxidation state of carbon by using rule 3, which states that the sum of the oxidation states of all the atoms must be 0.	(d) CO_2 (C ox state) + 2(O ox state) = 0 (C ox state) + 2(−2) = 0 C ox state = +4 $\underset{+4}{C}\underset{-2}{O_2}$ sum: $+4+2(-2)=0$
The oxidation state of oxygen is −2 (rule 5). You would ordinarily expect the oxidation state of S to be −2 (rule 5). However, if that were the case, the sum of the oxidation states would not equal the charge of the ion. Since O is higher on the list than S, it takes priority. Deduce the oxidation state of sulfur by setting the *sum* of all of the oxidation states equal to −2 (the charge of the ion).	(e) SO_4^{2-} (S ox state) + 4(O ox state) = −2 (S ox state) + 4(−2) = −2 S ox state = +6 $\underset{+6}{S}\underset{-2}{O_4}{}^{2-}$ sum: $+6+4(-2)=-2$
The oxidation state of potassium is +1 (rule 4). You would ordinarily expect the oxidation state of O to be −2 (rule 5), but rule 4 takes priority. Deduce the oxidation state of O by setting the sum of all of the oxidation states equal to 0.	(f) K_2O_2 2(K ox state) + 2(O ox state) = 0 2(+1) + 2(O ox state) = 0 O ox state = −1 $\underset{+1}{K_2}\underset{-1}{O_2}$ sum: $2(+1)+2(-1)=0$

FOR PRACTICE 4.16

Assign an oxidation state to each atom in each species.

(a) Cr (b) Cr^{3+} (c) CCl_4 (d) $SrBr_2$ (e) SO_3 (f) NO_3^-

In most cases, oxidation states are positive or negative integers; however, on occasion an atom within a compound can have a fractional oxidation state. For example, consider KO_2. The oxidation state is assigned as follows:

$$\underset{+1}{K}\underset{-\frac{1}{2}}{O_2}$$

$$\text{sum: } +1+2\left(-\frac{1}{2}\right)=0$$

In KO_2, oxygen has a $-\frac{1}{2}$ oxidation state. Although this seems unusual, it is accepted because oxidation states are merely an imposed electron bookkeeping scheme, not an actual physical quantity.

Identifying Redox Reactions

We can use oxidation states to identify redox reactions, even between nonmetals. For example, is the following reaction between carbon and sulfur a redox reaction?

$$C + 2\,S \longrightarrow CS_2$$

If so, what element is oxidized? What element is reduced? We can use the oxidation state rules to assign oxidation states to all elements on both sides of the equation.

	C	+	$2\,S$	$\longrightarrow$	CS_2
Oxidation states:	0		0		+4 −2

Reduction

Oxidation

Carbon changed from an oxidation state of 0 to an oxidation state of +4. In terms of our electron bookkeeping scheme (the assigned oxidation state), carbon *lost electrons* and was *oxidized*. Sulfur changed from an oxidation state of 0 to an oxidation state of −2. In terms of our electron bookkeeping scheme, sulfur *gained electrons* and was *reduced*. In terms of oxidation states, oxidation and reduction are defined as follows:

- **Oxidation: An increase in oxidation state**
- **Reduction: A decrease in oxidation state**

Remember that reduction is a *reduction* in oxidation state.

EXAMPLE 4.17 Using Oxidation States to Identify Oxidation and Reduction

Use oxidation states to identify the element that is being oxidized and the element that is being reduced in the following redox reaction:

$$Mg(s) + 2\,H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

SOLUTION

Begin by assigning oxidation states to each atom in the reaction.

	$Mg(s)$	+	$2\,H_2O(l)$	$\longrightarrow$	$Mg(OH)_2(aq)$	+	$H_2(g)$
Oxidation states:	0		+1 −2		+2 −2 +1		0

Reduction

Oxidation

Since Mg increased in oxidation state, it was oxidized. Since H decreased in oxidation state, it was reduced.

FOR PRACTICE 4.17

Use oxidation states to identify the element that is being oxidized and the element that is being reduced in the following redox reaction:

$$Sn(s) + 4\,HNO_3(aq) \longrightarrow SnO_2(s) + 4\,NO_2(g) + 2\,H_2O(g)$$

FOR MORE PRACTICE 4.17

Determine if each reaction is a redox reaction. If the reaction is a redox reaction, identify which element is oxidized and which is reduced.

(a) $Hg_2(NO_3)_2(aq) + 2\,KBr(aq) \longrightarrow Hg_2Br_2(s) + 2\,KNO_3(aq)$

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

(c) $CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$

Notice that *oxidation and reduction must occur together*. If one substance loses electrons (oxidation), then another substance must gain electrons (reduction). A substance that causes the oxidation of another substance is an **oxidizing agent**.

Oxygen, for example, is an excellent oxidizing agent because it causes the oxidation of many other substances. In a redox reaction, *the oxidizing agent is always reduced*. A substance that causes the reduction of another substance is a **reducing agent**. Hydrogen, for example, as well as the group 1A and group 2A metals (because of their tendency to lose electrons) are excellent reducing agents. In a redox reaction, *the reducing agent is always oxidized*.

In Section 18.2 you will learn more about redox reactions, including how to balance them. For now, be able to identify redox reactions, as well as oxidizing and reducing agents, according to these guidelines.

Redox reactions include:

- **Any reaction in which there is a change in the oxidation states of atoms between the reactants and the products.**

In a redox reaction:

- **The oxidizing agent oxidizes another substance (and is itself reduced).**
- **The reducing agent reduces another substance (and is itself oxidized).**

EXAMPLE 4.18 Identifying Redox Reactions, Oxidizing Agents, and Reducing Agents

Determine whether each reaction is an oxidation–reduction reaction. If the reaction is an oxidation–reduction, identify the oxidizing agent and the reducing agent.

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

(b) $2\ HBr(aq) + Ca(OH)_2(aq) \rightarrow 2\ H_2O(l) + CaBr_2(aq)$

(c) $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$

SOLUTION

This is a redox reaction because magnesium increases in oxidation number (oxidation) and oxygen decreases in oxidation number (reduction).

(a)
$$\underset{0}{2\ Mg(s)} + \underset{0}{O_2(g)} \longrightarrow 2\ \underset{+2\ -2}{MgO(s)}$$

Reduction

Oxidation

Oxidizing agent: O_2
Reducing agent: Mg

This is not a redox reaction because none of the atoms undergoes a change in oxidation number.

(b)
$$2\ \underset{+1\ -1}{HBr(aq)} + \underset{+2\ -2\ +1}{Ca(OH)_2(aq)} \longrightarrow \underset{+1\ -2}{2H_2O(l)} + \underset{+2\ -1}{CaBr_2(aq)}$$

This is a redox reaction because zinc increases in oxidation number (oxidation) and iron decreases in oxidation number (reduction).

(c)
$$\underset{0}{Zn(s)} + \underset{+2}{Fe^{2+}(aq)} \longrightarrow \underset{+2}{Zn^{2+}(aq)} + \underset{0}{Fe(s)}$$

Reduction

Oxidation

Oxidizing agent: Fe^{2+}
Reducing agent: Zn

FOR PRACTICE 4.18

Determine whether each reaction is a redox reaction. For all redox reactions, identify the oxidizing agent and the reducing agent.

(a) $2\ Li(s) + Cl_2(g) \rightarrow 2\ LiCl(s)$

(b) $2\ Al(s) + 3\ Sn^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Sn(s)$

(c) $Pb(NO_3)_2(aq) + 2\ LiCl(aq) \rightarrow PbCl_2(s) + 2\ LiNO_3(aq)$

(d) $C(s) + O_2(g) \rightarrow CO_2(g)$

OXIDATION AND REDUCTION

CONCEPTUAL CONNECTION 4.8

Which statement is true regarding redox reactions?

(a) A redox reaction can occur without any changes in the oxidation states of the elements within the reactants and products of a reaction.

(b) If any of the reactants or products in a reaction contain oxygen, the reaction is a redox reaction.

(c) In a reaction, oxidation can occur independently of reduction.

(d) In a redox reaction, any increase in the oxidation state of a reactant must be accompanied by a decrease in the oxidation state of a reactant.

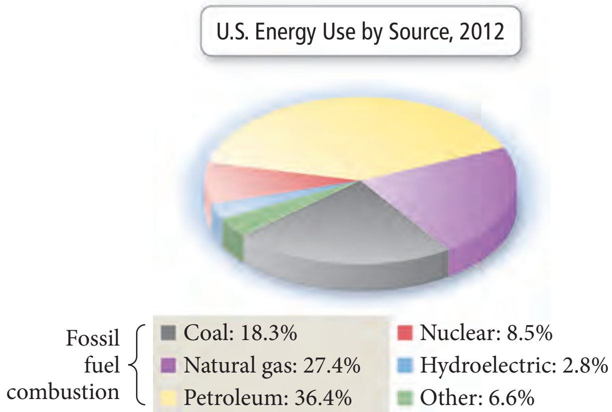
Combustion Reactions

We encountered combustion reactions, a type of redox reaction, in the opening section of this chapter. Combustion reactions are important because most of our society's energy is derived from them (FIGURE 4.19▸).

Combustion reactions are characterized by the reaction of a substance with O_2 to form one or more oxygen-containing compounds, often including water. Combustion reactions also emit heat. For example, as we saw earlier in this chapter, natural gas (CH_4) reacts with oxygen to form carbon dioxide and water.

$$\underset{-4\,+1}{CH_4(g)} + 2\,\underset{0}{O_2(g)} \longrightarrow \underset{+4\,-2}{CO_2(g)} + 2\,\underset{+1\,-2}{H_2O(g)}$$

Oxidation state: −4 +1 0 +4 −2 +1 −2



▲ **FIGURE 4.19 U.S. Energy Consumption** About 82% of the energy used in the United States in 2012 was produced by combustion reactions. Source: U.S. Energy Information Administration, *Monthly Energy Review*, December, 2013.

In this reaction, carbon is oxidized and oxygen is reduced. Ethanol, the alcohol in alcoholic beverages, also reacts with oxygen in a combustion reaction to form carbon dioxide and water.

$$C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$$

Compounds containing carbon and hydrogen—or carbon, hydrogen, and oxygen—always form carbon dioxide and water upon complete combustion. Other combustion reactions include the reaction of carbon with oxygen to form carbon dioxide

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

and the reaction of hydrogen with oxygen to form water.

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

EXAMPLE 4.19 Writing Equations for Combustion Reactions

Write a balanced equation for the combustion of liquid methyl alcohol (CH_3OH).

SOLUTION

Begin by writing an unbalanced equation showing the reaction of CH_3OH with O_2 to form CO_2 and H_2O.	$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$
Balance the equation applying the guidelines in Section 3.10.	$\mathbf{2}\,CH_3OH(l) + \mathbf{3}\,O_2(g) \longrightarrow \mathbf{2}\,CO_2(g) + \mathbf{4}\,H_2O(g)$

FOR PRACTICE 4.19

Write a balanced equation for the complete combustion of liquid C_2H_5SH.

Self-Assessment QUIZ

Q1. Manganese(IV) oxide reacts with aluminum to form elemental manganese and aluminum oxide.

$$3\ MnO_2 + 4\ Al \rightarrow 3\ Mn + 2\ Al_2O_3$$

What mass of Al is required to completely react with 25.0 g MnO_2?
a. 7.76 g Al
b. 5.82 g Al
c. 33.3 g Al
d. 10.3 g Al

Q2. Sodium and chlorine react to form sodium chloride.

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

What is the theoretical yield of sodium chloride when 55.0 g Na reacts with 67.2 g Cl_2?
a. 1.40×10^2 g NaCl
b. 111 g NaCl
c. 55.4 g NaCl
d. 222 g NaCl

Q3. Sulfur and fluorine react to form sulfur hexafluoride.

$$S(s) + 3\ F_2(g) \rightarrow SF_6(g)$$

If 50.0 g S is allowed to react as completely as possible with 105.0 g $F_2(g)$, what mass of excess reactant is left?
a. 20.5 g S
b. 45.7 g F_2
c. 15.0 g S
d. 36.3 g F_2

Q4. A reaction has a theoretical yield of 45.8 g. When the reaction is carried out, 37.2 g of the product is obtained. What is the percent yield?
a. 55.1%
b. 44.8%
c. 123%
d. 81.2%

Q5. What is the molarity of a solution containing 55.8 g of $MgCl_2$ dissolved in 1.00 L of solution?
a. 55.8 M
b. 1.71 M
c. 0.586 M
d. 0.558 M

Q6. What mass (in grams) of $Mg(NO_3)_2$ is present in 145 mL of a 0.150 M solution of $Mg(NO_3)_2$?
a. 3.23 g
b. 0.022 g
c. 1.88 g
d. 143 g

Q7. What volume of a 1.50 M HCl solution should you use to prepare 2.00 L of a 0.100 M HCl solution?
a. 0.300 L
b. 0.133 L
c. 30.0 L
d. 2.00 L

Q8. Potassium iodide reacts with lead(II) nitrate in the following precipitation reaction:

$$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + PbI_2(s)$$

What minimum volume of 0.200 M potassium iodide solution is required to completely precipitate all of the lead in 155.0 mL of a 0.112 M lead(II) nitrate solution?
a. 348 mL
b. 86.8 mL
c. 174 mL
d. 43.4 mL

Q9. Which solution will form a precipitate when mixed with a solution of aqueous Na_2CO_3?
a. $KNO_3(aq)$
b. $NaBr(aq)$
c. $NH_4Cl(aq)$
d. $CuCl_2(aq)$

Q10. What is the net ionic equation for the reaction that occurs when aqueous solutions of KOH and $SrCl_2$ are mixed?
a. $K^+(aq) + Cl^-(aq) \rightarrow KCl(s)$
b. $Sr^{2+}(aq) + 2\ OH^-(aq) \rightarrow Sr(OH)_2(s)$
c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
d. None of the above because no reaction occurs.

Q11. What is the net ionic equation for the reaction that occurs when aqueous solutions of KOH and HNO_3 are mixed?
a. $K^+(aq) + NO_3^-(aq) \rightarrow KNO_3(s)$
b. $NO_3^-(aq) + OH^-(aq) \rightarrow NO_3OH(s)$
c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
d. None of the above because no reaction occurs.

Q12. What is the net ionic equation for the reaction that occurs when aqueous solutions of $KHCO_3$ and HBr are mixed?
a. $K^+(aq) + Br^-(aq) \rightarrow KBr(s)$
b. $H^+(aq) + HCO_3^-(aq) \rightarrow CO_2(g) + H_2O(l)$
c. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
d. None of the above because no reaction occurs.

Q13. What is the oxidation state of carbon in CO_3^{2-}?
a. +4 **b.** +3 **c.** −3 **d.** −2

Q14. Sodium reacts with water according to the following reaction:

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

Identify the oxidizing agent.
a. $Na(s)$
b. $H_2O(l)$
c. $NaOH(aq)$
d. $H_2(aq)$

Q15. Identify the correct balanced equation for the combustion of propane (C_3H_8).
a. $C_3H_8(g) \rightarrow 4\ H_2(g) + 3\ C(s)$
b. $C_3H_8(g) + 5\ O_2(g) \rightarrow 4\ H_2O(g) + 3\ CO_2(g)$
c. $C_3H_8(g) + 3\ O_2(g) \rightarrow 4\ H_2O(g) + 3\ CO_2(g)$
d. $2\ C_3H_8(g) + 9\ O_2(g) \rightarrow 6\ H_2CO_3(g) + 2\ H_2(g)$

Answers: 1:d; 2:b; 3:a; 4:d; 5:c; 6:a; 7:b; 8:c; 9:d; 10:b; 11:c; 12:b; 13:a; 14:b; 15:b

Chapter 4 in Review

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Key Concepts

Climate Change and the Combustion of Fossil Fuels (4.1)
- Climate change, caused by rising atmospheric carbon dioxide levels, is potentially harmful. We can use reaction stoichiometry to verify that the largest carbon dioxide source is the burning of fossil fuels.

Reaction Stoichiometry (4.2)
- Reaction stoichiometry refers to the numerical relationships between the reactants and products in a balanced chemical equation.
- Reaction stoichiometry allows us to predict, for example, the amount of product that can be formed for a given amount of reactant, or how much of one reactant is required to react with a given amount of another.

Limiting Reactant, Theoretical Yield, and Percent Yield (4.3)
- The limiting reactant in a chemical reaction is the reactant that is present in the smallest stoichiometric quantity; it will be completely consumed in the reaction, and it limits the amount of product that can be made.
- Any reactant that does not limit the amount of product is said to be in excess.
- The amount of product that can be made from the limiting reactant is the theoretical yield.
- The actual yield—always equal to or less than the theoretical yield—is the amount of product that is actually made when the reaction is carried out.
- The percentage of the theoretical yield that is actually produced is the percent yield.

Solution Concentration and Stoichiometry (4.4)
- An aqueous solution is a homogeneous mixture of water (the solvent) with another substance (the solute).
- The concentration of a solution is often expressed in molarity, the number of moles of solute per liter of solution.
- We can use the molarities and volumes of reactant solutions to predict the amount of product that will form in an aqueous reaction.

Aqueous Solutions and Precipitation Reactions (4.5, 4.6)
- Solutes that completely dissociate (or completely ionize in the case of the acids) to ions in solution are strong electrolytes and are good conductors of electricity.
- Solutes that only partially dissociate (or partially ionize) are weak electrolytes, and solutes that do not dissociate (or ionize) at all are nonelectrolytes.
- A substance that dissolves in water to form a solution is said to be soluble.
- In a precipitation reaction, two aqueous solutions are mixed and a solid—or precipitate—forms.
- The solubility rules are empirical guidelines that help predict the solubilities of ionic compounds; these rules are especially useful when determining whether or not a precipitate will form.

Equations for Aqueous Reactions (4.7)
- We can represent an aqueous reaction with a molecular equation, which shows the complete neutral formula for each compound in the reaction.
- An aqueous reaction can also be represented with a complete ionic equation, which shows the dissociated nature of the aqueous ionic compounds.
- A third representation of an aqueous reaction is a net ionic equation, in which the spectator ions—those that do not change in the course of the reaction—are left out of the equation.

Acid–Base and Gas-Evolution Reactions (4.8)
- In an acid–base reaction, an acid, a substance that produces H^+ in solution, reacts with a base, a substance that produces OH^- in solution, and the two neutralize each other, producing water (or in some cases a weak electrolyte).
- In gas-evolution reactions, two aqueous solutions are combined and a gas is produced.

Oxidation–Reduction Reactions (4.9)
- In oxidation–reduction reactions, one substance transfers electrons to another substance.

- The substance that loses electrons is oxidized and the one that gains them is reduced.
- An oxidation state is a charge given to each atom in a redox reaction by assigning all shared electrons to the atom with the greater attraction for those electrons. Oxidation states are an imposed bookkeeping scheme, not actual physical states.
- The oxidation state of an atom increases upon oxidation and decreases upon reduction.
- A combustion reaction is a specific type of oxidation—a reduction reaction in which a substance reacts with oxygen—emitting heat and forming one or more oxygen-containing products.

Key Equations and Relationships

Mass-to-Mass Conversion: Stoichiometry (4.2)

mass A $\longrightarrow$ amount A (in moles) $\longrightarrow$ amount B (in moles) $\longrightarrow$ mass B

Percent Yield (4.3)

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

Molarity (M): Solution Concentration (4.4)

$$\text{M} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

Solution Dilution (4.4)

$$M_1V_1 = M_2V_2$$

Solution Stoichiometry (4.4)

volume A $\longrightarrow$ amount A (in moles) $\longrightarrow$ amount B (in moles) $\longrightarrow$ volume B

Key Learning Objectives

Chapter Objectives	Assessment
Calculating Stoichiometric Quantities (4.2)	Examples 4.1, 4.2 For Practice 4.1, 4.2 Exercises 7–12
Determining the Limiting Reactant and Calculating Theoretical and Percent Yield (4.3)	Examples 4.3, 4.4 For Practice 4.3, 4.4 Exercises 19–24
Calculating and Using Molarity as a Conversion Factor (4.4)	Examples 4.5, 4.6 For Practice 4.5, 4.6 For More Practice 4.5, 4.6 Exercises 25–32
Determining Solution Dilutions (4.4)	Example 4.7 For Practice 4.7 For More Practice 4.7 Exercises 33–36
Using Solution Stoichiometry to Find Volumes and Amounts (4.4)	Example 4.8 For Practice 4.8 For More Practice 4.8 Exercises 37–40
Predicting Whether a Compound Is Soluble (4.5)	Example 4.9 For Practice 4.9 Exercises 45, 46
Writing Equations for Precipitation Reactions (4.6)	Examples 4.10, 4.11 For Practice 4.10, 4.11 Exercises 47–50
Writing Complete Ionic and Net Ionic Equations (4.7)	Example 4.12 For Practice 4.12 For More Practice 4.12 Exercises 51–54
Writing Equations for Acid–Base Reactions (4.8)	Examples 4.13, 4.14 For Practice 4.13, 4.14 Exercises 55–60
Writing Equations for Gas-Evolution Reactions (4.8)	Example 4.15 For Practice 4.15 For More Practice 4.15 Exercises 61, 62
Assigning Oxidation States (4.9)	Example 4.16 For Practice 4.16 Exercises 63–66
Identifying Redox Reactions, Oxidizing Agents, and Reducing Agents Using Oxidation States (4.9)	Examples 4.17, 4.18 For Practice 4.17, 4.18 For More Practice 4.17 Exercises 67, 68
Writing Equations for Combustion Reactions (4.9)	Example 4.19 For Practice 4.19 Exercises 69, 70

Exercises

Problems by Topic

Reaction Stoichiometry

1. Consider the unbalanced equation for the combustion of hexane.
$$C_6H_{14}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$
Balance the equation and determine how many moles of O_2 are required to react completely with 4.9 moles C_6H_{14}.

2. Consider the unbalanced equation for the neutralization of acetic acid.
$$HC_2H_3O_2(aq) + Ba(OH)_2(aq) \longrightarrow H_2O(l) + Ba(C_2H_3O_2)_2(aq)$$
Balance the equation and determine how many moles of $Ba(OH)_2$ are required to completely neutralize 0.107 moles of $HC_2H_3O_2$.

3. For the reaction shown, calculate how many moles of NO_2 form when each amount of reactant completely reacts.

$$2\,N_2O_5(g) \rightarrow 4\,NO_2(g) + O_2(g)$$

a. 1.3 mol N_2O_5 **b.** 5.8 mol N_2O_5
c. 10.5 g N_2O_5 **d.** 1.55 kg N_2O_5

4. For the reaction shown, calculate how many moles of NH_3 form when each amount of reactant completely reacts.

$$3\,N_2H_4(l) \rightarrow 4\,NH_3(g) + N_2(g)$$

a. 5.3 mol N_2H_4 **b.** 2.28 mol N_2H_4
c. 32.5 g N_2H_4 **d.** 14.7 kg N_2H_4

5. Consider the balanced equation.

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

Complete the table showing the appropriate number of moles of reactants and products. When the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product formed. When the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that is made.

Mol SiO_2	Mol C	Mol SiC	Mol CO
3	_____	_____	_____
_____	6	_____	_____
_____	_____	_____	10
2.8	_____	_____	_____
_____	1.55	_____	_____

6. Consider the balanced equation.

$$2\,N_2H_4(g) + N_2O_4(g) \rightarrow 3\,N_2(g) + 4\,H_2O(g)$$

Complete the table showing the appropriate number of moles of reactants and products. When the number of moles of a *reactant* is provided, fill in the required amount of the other reactant, as well as the moles of each product formed. When the number of moles of a *product* is provided, fill in the required amount of each reactant to make that amount of product, as well as the amount of the other product that is made.

Mol N_2H_4	Mol N_2O_4	Mol N_2	Mol H_2O
2	_____	_____	_____
_____	5	_____	_____
_____	_____	_____	10
2.5	_____	_____	_____
_____	4.2	_____	_____
_____	_____	11.8	_____

7. Hydrobromic acid dissolves solid iron according to this reaction.

$$Fe(s) + 2\,HBr(aq) \rightarrow FeBr_2(aq) + H_2(g)$$

What mass of HBr (in g) would dissolve a 4.8-g pure iron bar on a padlock? What mass of H_2 would be produced by the complete reaction of the iron bar?

8. Sulfuric acid dissolves aluminum metal according to the reaction shown here.

$$2\,Al(s) + 3\,H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$$

Suppose you wanted to dissolve an aluminum block with a mass of 12.7 g. What minimum mass of H_2SO_4 (in g) would you need? What mass of H_2 gas (in g) would be produced by the complete reaction of the aluminum block?

9. For each reaction, calculate the mass (in grams) of the product that forms when 2.5 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

a. $\underline{Ba(s)} + Cl_2(g) \rightarrow BaCl_2(s)$
b. $\underline{CaO(s)} + CO_2(g) \rightarrow CaCO_3(s)$
c. $2\,\underline{Mg(s)} + O_2(g) \rightarrow 2\,MgO(s)$
d. $4\,\underline{Al(s)} + 3\,O_2(g) \rightarrow 2\,Al_2O_3(s)$

10. For each reaction, calculate the mass (in grams) of the product that forms when 10.4 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

a. $2\,K(s) + \underline{Cl_2(g)} \rightarrow 2\,KCl(s)$
b. $2\,K(s) + \underline{Br_2(l)} \rightarrow 2\,KBr(s)$
c. $4\,Cr(s) + 3\,\underline{O_2(g)} \rightarrow 2\,Cr_2O_3(s)$
d. $2\,\underline{Sr(s)} + O_2(g) \rightarrow 2\,SrO(s)$

11. For each acid–base reaction, calculate the mass (in grams) of each acid necessary to completely react with and neutralize 4.85 g of the base.

a. $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$
b. $2\,HNO_3(aq) + Ca(OH)_2(aq) \rightarrow 2\,H_2O(l) + Ca(NO_3)_2(aq)$
c. $H_2SO_4(aq) + 2\,KOH(aq) \rightarrow 2\,H_2O(l) + K_2SO_4(aq)$

12. For each precipitation reaction, calculate how many grams of the first reactant are necessary to completely react with 55.8 g of the second reactant.

a. $2\,KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$
b. $Na_2CO_3(aq) + CuCl_2(aq) \rightarrow CuCO_3(s) + 2\,NaCl(aq)$
c. $K_2SO_4(aq) + Sr(NO_3)_2(aq) \rightarrow SrSO_4(s) + 2\,KNO_3(aq)$

Limiting Reactant, Theoretical Yield, and Percent Yield

13. For the reaction shown, find the limiting reactant for each of the initial amounts of reactants.

$$2\,Na(s) + Br_2(g) \rightarrow 2\,NaBr(s)$$

a. 2 mol Na, 2 mol Br_2
b. 1.8 mol Na, 1.4 mol Br_2
c. 2.5 mol Na, 1 mol Br_2
d. 12.6 mol Na, 6.9 mol Br_2

14. For the reaction shown, find the limiting reactant for each of the initial amounts of reactants.

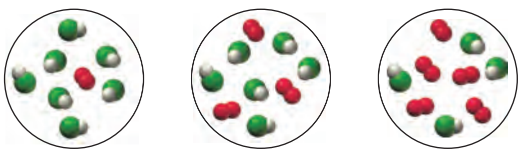
$$4\,Al(s) + 3\,O_2(g) \rightarrow 2\,Al_2O_3(s)$$

a. 1 mol Al, 1 mol O_2
b. 4 mol A1, 2.6 mol O_2
c. 16 mol Al, 13 mol O_2
d. 7.4 mol Al, 6.5 mol O_2

15. Consider the reaction.

$$4\,HCl(g) + O_2(g) \rightarrow 2\,H_2O(g) + 2\,Cl_2(g)$$

Each molecular diagram represents an initial mixture of the reactants. How many molecules of Cl_2 are formed from the reaction mixture that produces the greatest amount of products?



a. **b.** **c.**

16. Consider the reaction.

$$2\,CH_3OH(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

Each molecular diagram represents an initial mixture of the reactants. How many CO_2 molecules are formed from the reaction mixture that produces the greatest amount of products?

a. **b.** **c.**

17. For the reaction shown, calculate the theoretical yield of the product (in moles) for each set of initial amounts of reactants.

$$Ti(s) + 2\,Cl_2(g) \rightarrow TiCl_4(l)$$

a. 4 mol Ti, 4 mol Cl_2
b. 7 mol Ti, 17 mol Cl_2
c. 12.4 mol Ti, 18.8 mol Cl_2

18. For the reaction shown, calculate the theoretical yield of product (in moles) for each set of initial amounts of reactants.

$$2\,Mn(s) + 2\,O_2(g) \rightarrow 2\,MnO_2(s)$$

a. 3 mol Mn, 3 mol O_2 **b.** 4 mol Mn, 7 mol O_2
c. 27.5 mol Mn, 43.8 mol O_2

19. For the reaction shown, calculate the theoretical yield of product (in grams) for each set of the initial amounts of reactants.

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

a. 2.0 g Al, 2.0 g Cl_2 **b.** 7.5 g Al, 24.8 g Cl_2
c. 0.235 g Al, 1.15 g Cl_2

20. For the reaction shown, calculate the theoretical yield of the product (in grams) for each set of the initial amounts of reactants.

$$Ti(s) + 2\,F_2(g) \rightarrow TiF_4(s)$$

a. 5.0 g Ti, 5.0 g F_2 **b.** 2.4 g Ti, 1.6 g F_2
c. 0.233 g Ti, 0.288 g F_2

21. Lead ions can be precipitated from solution with KCl according to the reaction:

$$Pb^{2+}(aq) + 2\,KCl(aq) \rightarrow PbCl_2(s) + 2\,K^+(aq)$$

When 28.5 g KCl is added to a solution containing 25.7 g Pb^{2+}, a $PbCl_2$ precipitate forms. The precipitate is filtered, dried, and found to have a mass of 29.4 g. Determine the limiting reactant, theoretical yield of $PbCl_2$, and percent yield for the reaction.

22. Magnesium oxide can be made by heating magnesium metal in the presence of oxygen. The balanced equation for the reaction is shown here.

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

When 10.1 g of Mg reacts with 10.5 g O_2, 11.9 g MgO is collected. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

23. Urea (CH_4N_2O) is a common fertilizer that can be synthesized by the reaction of ammonia (NH_3) with carbon dioxide as follows:

$$2\,NH_3(aq) + CO_2(aq) \rightarrow CH_4N_2O(aq) + H_2O(l)$$

In an industrial synthesis of urea, a chemist combines 136.4 kg of ammonia with 211.4 kg of carbon dioxide and obtains 168.4 kg of urea. Determine the limiting reactant, theoretical yield of urea, and percent yield for the reaction.

24. Many computer chips are manufactured from silicon, which occurs in nature as SiO_2. When SiO_2 is heated to melting, it reacts with solid carbon to form liquid silicon and carbon monoxide gas. In an industrial preparation of silicon, 155.8 kg of SiO_2 reacts with 78.3 kg of carbon to produce 66.1 kg of silicon. Determine the limiting reactant, theoretical yield, and percent yield for the reaction.

Solution Concentration and Solution Stoichiometry

25. Calculate the molarity of each solution.
a. 4.3 mol of LiCl in 2.8 L solution
b. 22.6 g $C_6H_{12}O_6$ in 1.08 L of solution
c. 45.5 mg NaCl in 154.4 mL of solution

26. Calculate the molarity of each solution.
a. 0.11 mol of $LiNO_3$ in 5.2 L of solution
b. 61.3 g C_2H_6O in 2.44 L of solution
c. 15.2 mg KI in 102 mL of solution

27. What is the molarity of NO_3^- in each solution?
a. 0.150 M KNO_3 **b.** 0.150 M $Ca(NO_3)_2$
c. 0.150 M $Al(NO_3)_3$

28. What is the molarity of Cl^- in each solution?
a. 0.200 M NaCl **b.** 0.150 M $SrCl_2$
c. 0.100 M $AlCl_3$

29. How many moles of KCl are in each solution?
a. 0.556 L of a 2.3 M KCl solution
b. 1.8 L of a 0.85 M KCl solution
c. 114 mL of a 1.85 M KCl solution

30. What volume of 0.200 M ethanol solution contains each of the number of moles of ethanol?
a. 0.45 mol ethanol
b. 1.22 mol ethanol
c. 1.2×10^{-2} mol ethanol

31. A laboratory procedure calls for making 500.0 mL of a 1.3 M $NaNO_3$ solution. What mass of $NaNO_3$ (in g) is needed?

32. A chemist wants to make 7.2 L of a 0.350 M $CaCl_2$ solution. What mass of $CaCl_2$ (in g) should the chemist use?

33. If 123 mL of a 1.1 M glucose solution is diluted to 500.0 mL, what is the molarity of the diluted solution?

34. If 3.5 L of a 4.8 M $SrCl_2$ solution is diluted to 45 L, what is the molarity of the diluted solution?

35. To what volume should you dilute 50.0 mL of a 12 M stock HNO_3 solution to obtain a 0.100 M HNO_3 solution?

36. To what volume should you dilute 25 mL of a 10.0 M H_2SO_4 solution to obtain a 0.150 M H_2SO_4 solution?

37. Consider the precipitation reaction.

$$2\,Na_3PO_4(aq) + 3\,CuCl_2(aq) \rightarrow Cu_3(PO_4)_2(s) + 6\,NaCl(aq)$$

What volume of 0.175 M Na_3PO_4 solution is necessary to completely react with 95.4 mL of 0.102 M $CuCl_2$?

38. Consider the precipitation reaction.

$$Li_2S(aq) + Co(NO_3)_2(aq) \rightarrow 2\,LiNO_3(aq) + CoS(s)$$

What volume of 0.150 M Li_2S solution is required to completely react with 125 mL of 0.150 M $Co(NO_3)_2$?

39. What is the minimum amount of 6.0 M H_2SO_4 necessary to produce 25.0 g of $H_2(g)$ according to the following reaction?

$$2\,Al(s) + 3\,H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3\,H_2(g)$$

40. What is the molarity of $ZnCl_2$ that forms when 25.0 g of zinc completely reacts with $CuCl_2$ according to the following reaction? Assume a final volume of 275 mL.

$$Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$$

41. A 25.0 mL sample of a 1.20 M potassium chloride solution is mixed with 15.0 mL of a 0.900 M lead(II) nitrate solution, and the precipitation reaction shown here occurs.

$$2\ KCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ KNO_3(aq)$$

The solid $PbCl_2$ is collected, dried, and found to have a mass of 2.45 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

42. A 55.0 mL sample of a 0.102 M potassium sulfate solution is mixed with 35.0 mL of a 0.114 M lead(II) acetate solution and the precipitation reaction shown here occurs.

$$K_2SO_4(aq) + Pb(C_2H_3O_2)_2(aq) \rightarrow 2\ KC_2H_3O_2(aq) + PbSO_4(s)$$

The solid $PbSO_4$ is collected, dried, and found to have a mass of 1.01 g. Determine the limiting reactant, the theoretical yield, and the percent yield.

Types of Aqueous Solutions and Solubility

43. Each compound listed is soluble in water. For each compound, do you expect the resulting aqueous solution to conduct electrical current?
a. CsCl **b.** CH_3OH **c.** $Ca(NO_2)_2$ **d.** $C_6H_{12}O_6$

44. Classify each compound as a strong electrolyte or nonelectrolyte.
a. $MgBr_2$ **b.** $C_{12}H_{22}O_{11}$ **c.** Na_2CO_3 **d.** KOH

45. Determine whether each compound is soluble or insoluble. If the compound is soluble, write the ions present in solution.
a. $AgNO_3$ **b.** $Pb(C_2H_3O_2)_2$
c. KNO_3 **d.** $(NH_4)_2S$

46. Determine whether each compound is soluble or insoluble. For the soluble compounds, write the ions present in solution.
a. AgI **b.** $Cu_3(PO_4)_2$ **c.** $CoCO_3$ **d.** K_3PO_4

Precipitation Reactions

47. Complete and balance each equation. If no reaction occurs, write NO REACTION.
a. $LiI(aq) + BaS(aq) \rightarrow$
b. $KCl(aq) + CaS(aq) \rightarrow$
c. $CrBr_2(aq) + Na_2CO_3(aq) \rightarrow$
d. $NaOH(aq) + FeCl_3(aq) \rightarrow$

48. Complete and balance each equation. If no reaction occurs, write NO REACTION.
a. $NaNO_3(aq) + KCl(aq) \rightarrow$
b. $NaCl(aq) + Hg_2(C_2H_3O_2)_2(aq) \rightarrow$
c. $(NH_4)_2SO_4(aq) + SrCl_2(aq) \rightarrow$
d. $NH_4Cl(aq) + AgNO_3(aq) \rightarrow$

49. Write a molecular equation for the precipitation reaction that occurs (if any) when each set of solutions is mixed. If no reaction occurs, write NO REACTION.
a. potassium carbonate and lead(II) nitrate
b. lithium sulfate and lead(II) acetate
c. copper(II) nitrate and magnesium sulfide
d. strontium nitrate and potassium iodide

50. Write a molecular equation for the precipitation reaction that occurs (if any) when each set of solutions is mixed. If no reaction occurs, write NO REACTION.
a. sodium chloride and lead(II) acetate
b. potassium sulfate and strontium iodide
c. cesium chloride and calcium sulfide
d. chromium(III) nitrate and sodium phosphate

Ionic and Net Ionic Equations

51. Write balanced complete ionic and net ionic equations for each reaction.
a. $HCl(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCl(aq)$
b. $CaS(aq) + CuCl_2(aq) \rightarrow CuS(s) + CaCl_2(aq)$
c. $NaOH(aq) + HNO_3(aq) \rightarrow H_2O(l) + NaNO_3(aq)$
d. $Na_3PO_4(aq) + NiCl_2(aq) \rightarrow Ni_3(PO_4)_2(s) + NaCl(aq)$

52. Write balanced complete ionic and net ionic equations for each reaction.
a. $K_2SO_4(aq) + CaI_2(aq) \rightarrow CaSO_4(s) + KI(aq)$
b. $NH_4Cl(aq) + NaOH(aq) \rightarrow H_2O(l) + NH_3(g) + NaCl(aq)$
c. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
d. $HC_2H_3O_2(aq) + K_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + KC_2H_3O_2(aq)$

53. Mercury ions (Hg_2^{2+}) can be removed from solution by precipitation with Cl^-. Suppose that a solution contains aqueous $Hg_2(NO_3)_2$. Write complete ionic and net ionic equations to show the reaction of aqueous $Hg_2(NO_3)_2$ with aqueous sodium chloride to form solid Hg_2Cl_2 and aqueous sodium nitrate.

54. Lead ions can be removed from solution by precipitation with sulfate ions. Suppose that a solution contains lead(II) nitrate. Write complete ionic and net ionic equations to show the reaction of aqueous lead(II) nitrate with aqueous potassium sulfate to form solid lead(II) sulfate and aqueous potassium nitrate.

Acid–Base and Gas-Evolution Reactions

55. Write balanced molecular and net ionic equations for the reaction between hydrobromic acid and potassium hydroxide.

56. Write balanced molecular and net ionic equations for the reaction between nitric acid and calcium hydroxide.

57. Complete and balance each acid–base reaction equation.
a. $H_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow$
b. $HClO_4(aq) + KOH(aq) \rightarrow$
c. $H_2SO_4(aq) + NaOH(aq) \rightarrow$

58. Complete and balance each acid–base reaction equation.
a. $HI(aq) + LiOH(aq) \rightarrow$
b. $HC_2H_3O_2(aq) + Ca(OH)_2(aq) \rightarrow$
c. $HCl(aq) + Ba(OH)_2(aq) \rightarrow$

59. Write balanced complete ionic and net ionic equations for each acid–base reaction.
a. $HBr(aq) + NaOH(aq) \rightarrow$
b. $HF(aq) + NaOH(aq) \rightarrow$
c. $HC_2H_3O_2(aq) + RbOH(aq) \rightarrow$

60. Write balanced complete ionic and net ionic equations for each acid–base reaction.
a. $HI(aq) + RbOH(aq) \rightarrow$
b. $HCHO_2(aq) + NaOH(aq) \rightarrow$
c. $HC_2H_3O_2(aq) + LiOH(aq) \rightarrow$

61. Complete and balance each gas-evolution reaction equation.
a. $HBr(aq) + NiS(s) \rightarrow$
b. $NH_4I(aq) + NaOH(aq) \rightarrow$
c. $HBr(aq) + Na_2S(aq) \rightarrow$
d. $HClO_4(aq) + Li_2CO_3(aq) \rightarrow$

62. Complete and balance each gas-evolution reaction equation.
a. $HNO_3(aq) + Na_2SO_3(aq) \rightarrow$
b. $HCl(aq) + KHCO_3(aq) \rightarrow$
c. $HC_2H_3O_2(aq) + NaHSO_3(aq) \rightarrow$
d. $(NH_4)_2SO_4(aq) + Ca(OH)_2(aq) \rightarrow$

Oxidation–Reduction and Combustion

63. Assign oxidation states to each atom in each ion or compound.
a. Ag **b.** Ag^+ **c.** CaF_2
d. H_2S **e.** CO_3^{2-} **f.** CrO_4^{2-}

64. Assign oxidation states to each atom in each ion or compound.
a. Cl_2 **b.** Fe^{3+} **c.** $CuCl_2$
d. CH_4 **e.** $Cr_2O_7^{2-}$ **f.** HSO_4^-

65. What is the oxidation state of Cr in each compound?
a. CrO **b.** CrO_3 **c.** Cr_2O_3

66. What is the oxidation state of Cl in each ion?
a. ClO^- **b.** ClO_2^- **c.** ClO_3^- **d.** ClO_4^-

67. Indicate which reactions are redox reactions. For each redox reaction, identify the oxidizing agent and the reducing agent.
a. $4\,Li(s) + O_2(g) \rightarrow 2\,Li_2O(s)$
b. $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$
c. $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2\,NaNO_3(aq)$
d. $HBr(aq) + KOH(aq) \rightarrow H_2O(l) + KBr(aq)$

68. Indicate which reactions are redox reactions. For each redox reaction, identify the oxidizing agent and the reducing agent.
a. $Al(s) + 3\,Ag^+(aq) \rightarrow Al^{3+}(aq) + 3\,Ag(s)$
b. $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
c. $Ba(s) + Cl_2(g) \rightarrow BaCl_2(s)$
d. $Mg(s) + Br_2(l) \rightarrow MgBr_2(s)$

69. Complete and balance each combustion reaction equation.
a. $S(s) + O_2(g) \rightarrow$ **b.** $C_3H_6(g) + O_2(g) \rightarrow$
c. $Ca(s) + O_2(g) \rightarrow$ **d.** $C_5H_{12}S(l) + O_2(g) \rightarrow$

70. Complete and balance each combustion reaction equation.
a. $C_4H_6(g) + O_2(g) \rightarrow$ **b.** $C(s) + O_2(g) \rightarrow$
c. $CS_2(l) + O_2(g) \rightarrow$ **d.** $C_3H_8O(l) + O_2(g) \rightarrow$

Cumulative Problems

71. The density of a 20.0% by mass ethylene glycol ($C_2H_6O_2$) solution in water is 1.03 g/mL. Find the molarity of the solution.

72. Find the percent by mass of sodium chloride in a 1.35 M NaCl solution. The density of the solution is 1.05 g/mL.

73. Sodium bicarbonate is often used to neutralize excess hydrochloric acid in an upset stomach. What mass of hydrochloric acid (in grams) can be neutralized by 2.5 g of sodium bicarbonate? (Hint: Begin by writing a balanced equation for the reaction between aqueous sodium bicarbonate and aqueous hydrochloric acid.)

74. Toilet bowl cleaners often contain hydrochloric acid to dissolve the calcium carbonate deposits that accumulate within a toilet bowl. What mass of calcium carbonate (in grams) can be dissolved by 3.8 g of HCl? (Hint: Begin by writing a balanced equation for the reaction between hydrochloric acid and calcium carbonate.)

75. The combustion of gasoline produces carbon dioxide and water. Assume gasoline to be pure octane (C_8H_{18}) and calculate the mass (in kg) of carbon dioxide that is added to the atmosphere per 1.0 kg of octane burned. (Hint: Begin by writing a balanced equation for the combustion reaction.)

76. Many home barbeques are fueled with propane gas (C_3H_8). What mass of carbon dioxide (in kg) is produced upon the complete combustion of 18.9 L of propane (approximate contents of one 5-gallon tank)? Assume that the density of the liquid propane in the tank is 0.621 g/mL. (Hint: Begin by writing a balanced equation for the combustion reaction.)

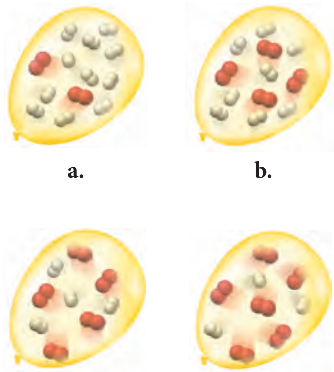
77. Aspirin can be made in the laboratory by reacting acetic anhydride ($C_4H_6O_3$) with salicylic acid ($C_7H_6O_3$) to form aspirin ($C_9H_8O_4$) and acetic acid ($HC_2H_3O_2$). The balanced equation is shown here.

$$C_4H_6O_3 + C_7H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$$

In a laboratory synthesis, a student begins with 3.00 mL of acetic anhydride (density = 1.08 g/mL) and 1.25 g of salicylic acid. Once the reaction is complete, the student collects 1.22 g of aspirin. Determine the limiting reactant, theoretical yield of aspirin, and percent yield for the reaction.

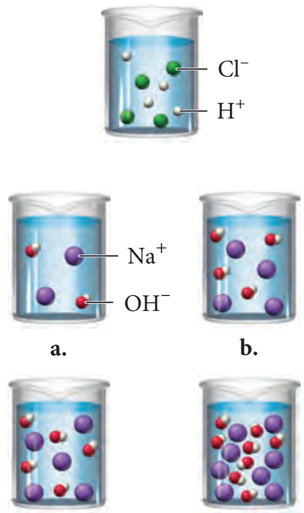
78. The combustion of liquid ethanol (C_2H_5OH) produces carbon dioxide and water. After 4.62 mL of ethanol (density = 0.789 g/mL) is allowed to burn in the presence of 15.55 g of oxygen gas, 3.72 mL of water (density = 1.00 g/mL) is collected. Determine the limiting reactant, theoretical yield of H_2O, and percent yield for the reaction. (Hint: Write a balanced equation for the combustion of ethanol.)

79. A loud classroom demonstration involves igniting a hydrogen-filled balloon. The hydrogen within the balloon reacts explosively with oxygen in the air to form water. If the balloon is filled with a mixture of hydrogen and oxygen, the explosion is even louder than if the balloon is filled with only hydrogen; the intensity of the explosion depends on the relative amounts of oxygen and hydrogen within the balloon. Look at the molecular views shown here, which represent different amounts of hydrogen and oxygen in four different balloons. Based on the balanced chemical equation, which balloon will make the loudest explosion?



a. b. c. d.

80. A hydrochloric acid solution will neutralize a sodium hydroxide solution. Look at the molecular views showing one beaker of HCl and four beakers of NaOH. Which NaOH beaker will just neutralize the HCl beaker? Begin by writing a balanced chemical equation for the neutralization reaction.



a. b. c. d.

81. Predict the products of each reaction and write balanced molecular equations for each. If no reaction occurs, write NO REACTION.
 a. $HCl(aq) + Hg_2(NO_3)_2(aq) \rightarrow$
 b. $KHSO_3(aq) + HNO_3(aq) \rightarrow$
 c. aqueous ammonium chloride and aqueous lead(II) nitrate
 d. aqueous ammonium chloride and aqueous calcium hydroxide

82. Predict the products of each of these reactions and write balanced molecular equations for each. If no reaction occurs, write NO REACTION.
 a. $H_2SO_4(aq) + HNO_3(aq) \rightarrow$
 b. $Cr(NO_3)_3(aq) + LiOH(aq) \rightarrow$
 c. liquid pentanol ($C_5H_{12}O$) and gaseous oxygen
 d. aqueous strontium sulfide and aqueous copper(II) sulfate

83. Hard water often contains dissolved Ca^{2+} and Mg^{2+} ions. One way to soften water is to add phosphates. The phosphate ion forms insoluble precipitates with calcium and magnesium ions, removing them from solution. Suppose that a solution is 0.050 M in calcium chloride and 0.085 M in magnesium nitrate. What mass of sodium phosphate must be added to 1.5 L of this solution to completely eliminate the hard water ions? Assume complete reaction.

84. An acid solution is 0.100 M in HCl and 0.200 M in H_2SO_4. What volume of a 0.150 M KOH solution must be added to 500.0 mL of the acidic solution to completely neutralize all of the acid?

85. Find the mass of barium metal (in grams) that must react with O_2 to produce enough barium oxide to prepare 1.0 L of a 0.10 M solution of OH^-.

86. A solution contains Cr^{3+} ion and Mg^{2+} ion. The addition of 1.00 L of 1.51 M NaF solution is required to cause the complete precipitation of these ions as $CrF_3(s)$ and $MgF_2(s)$. The total mass of the precipitate is 49.6 g. Determine the mass of Cr^{3+} in the original solution.

87. The nitrogen in sodium nitrate and in ammonium sulfate is available to plants as fertilizer. Which is the more economical source of nitrogen, a fertilizer containing 30.0% sodium nitrate by weight and costing \$9.00 per 100 lb or one containing 20.0% ammonium sulfate by weight and costing \$8.10 per 100 lb?

88. Find the volume of 0.110 M hydrochloric acid necessary to react completely with 1.52 g $Al(OH)_3$.

89. Treatment of gold metal with BrF_3 and KF produces Br_2 and $KAuF_4$, a salt of gold. Identify the oxidizing agent and the reducing agent in this reaction. Find the mass of the gold salt that forms when a 73.5-g mixture of equal masses of all three reactants is prepared.

90. A solution is prepared by mixing 0.10 L of 0.12 M sodium chloride with 0.23 L of a 0.18 M $MgCl_2$ solution. What volume of a 0.20 M silver nitrate solution is required to precipitate all the Cl^- ion in the solution as AgCl?

91. A solution contains one or more of the following ions: Ag^+, Ca^{2+}, and Cu^{2+}. When sodium chloride is added to the solution, no precipitate forms. When sodium sulfate is added to the solution, a white precipitate forms. The precipitate is filtered off, and sodium carbonate is added to the remaining solution, producing a precipitate. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

92. A solution contains one or more of the following ions: Hg_2^{2+}, Ba^{2+}, and Fe^{2+}. When potassium chloride is added to the solution, a precipitate forms. The precipitate is filtered off, and potassium sulfate is added to the remaining solution and no precipitate forms. When potassium carbonate is added to the remaining solution, a precipitate forms. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

93. A liquid level mixture contains 30.35% hexane (C_6H_{14}) and 15.85% heptane (C_7H_{16}), and the rest is octane (C_8H_{18}). What maximum mass of carbon dioxide is produced by the complete combustion of 10.0 kg of this hydrocarbon mixture?

94. An ilmenite–sand mixture contains 22.8% ilmenite $FeTiO_3$ by mass, and the first reaction is carried out with a 90.8% yield. If the second reaction is carried out with an 85.9% yield, what mass of titanium can be obtained from 1.00 kg of the ilmenite–sand mixture?

Challenge Problems

95. Lakes that have been acidified by acid rain (HNO_3 and H_2SO_4) can be neutralized by a process called liming, in which limestone ($CaCO_3$) is added to the acidified water. What mass of limestone (in kg) would be required to completely neutralize a 15.2-billion-liter lake that is 1.8×10^{-5} M in H_2SO_4 and 8.7×10^{-6} M in HNO_3?

96. Recall from Section 4.6 that sodium carbonate is often added to laundry detergents to soften hard water and make the detergent more effective. Suppose that a particular detergent mixture is designed to soften hard water that is 3.5×10^{-3} M in Ca^{2+} and 1.1×10^{-3} M in Mg^{2+} and that the average capacity of a washing machine is 19.5 gallons of water. If the detergent requires using 0.65 kg detergent per load of laundry, determine what percentage (by mass) of the detergent must be sodium carbonate in order to completely precipitate all of the calcium and magnesium ions in an average load of laundry water.

97. Lead poisoning is a serious condition resulting from the ingestion of lead in food, water, or other environmental sources. It affects the central nervous system, leading to a variety of symptoms such as distractibility, lethargy, and loss of motor coordination. Lead poisoning is treated with chelating agents, substances that bind to metal ions, allowing the lead to be eliminated in the urine. A modern chelating agent used for this purpose is succimer ($C_4H_6O_4S_2$). Suppose you are trying to determine the appropriate dose for succimer treatment of lead poisoning. Assume that a patient's blood lead levels are 45 μg/dL, that total blood volume is 5.0 L, and that one mole of succimer binds one mole of lead. What minimum mass of succimer (in mg) is needed to bind all of the lead in the patient's bloodstream?

98. A particular kind of emergency breathing apparatus—often placed in mines, caves, or other places where oxygen might become depleted or where the air might become poisoned—works via the chemical reaction shown here.

$$4\ KO_2(s) + 2\ CO_2(g) \rightarrow 2\ K_2CO_3(s) + 3\ O_2(g)$$

Notice that the reaction produces O_2, which can be breathed, and absorbs CO_2, a product of respiration. Suppose you work for a company interested in producing a self-rescue breathing apparatus (based on the above reaction) that would allow the user to survive for 10 minutes in an emergency situation. What are the important chemical considerations in designing such a unit? Estimate how much KO_2 would be required for the apparatus. (Find any necessary additional information—such as human breathing rates—from appropriate sources. Assume that normal air is 20% oxygen.)

99. Metallic aluminum reacts with MnO_2 at elevated temperatures to form manganese metal and aluminum oxide. A mixture of the two reactants is 67.2 mole percent Al. Find the theoretical yield (in grams) of manganese from the reaction of 250 g of this mixture.

100. Hydrolysis of the compound B_5H_9 forms boric acid, H_3BO_3. Fusion of boric acid with sodium oxide forms a borate salt, $Na_2B_4O_7$. Without writing complete equations, find the mass (in grams) of B_5H_9 required to form 151 g of the borate salt by this reaction sequence.

Conceptual Problems

101. Consider the reaction.

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

The molar mass of K is 39.09 g/mol and that of O_2 is 32.00 g/mol. Without doing any calculations, pick the conditions under which potassium is the limiting reactant and explain your reasoning.

a. 170 g K, 31 g O_2 **b.** 16 g K, 2.5 g O_2
c. 165 kg K, 28 kg O_2 **d.** 1.5 g K, 0.38 g O_2

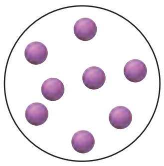
102. Consider the reaction.

$$2\,NO(g) + 5\,H_2(g) \longrightarrow 2\,NH_3(g) + 2\,H_2O(g)$$

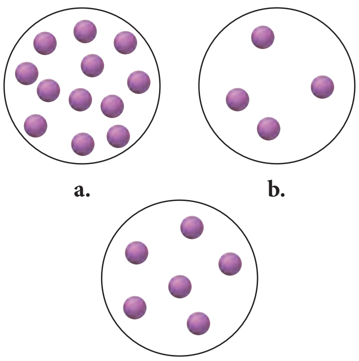
A reaction mixture initially contains 5 moles of NO and 10 moles of H_2. Without doing any calculations, determine which of the following best represents the mixture after the reactants have reacted as completely as possible. Explain your reasoning.

a. 1 mol NO, 0 mol H_2, 4 mol NH_3, 4 mol H_2O
b. 0 mol NO, 1 mol H_2, 5 mol NH_3, 5 mol H_2O
c. 3 mol NO, 5 mol H_2, 2 mol NH_3, 2 mol H_2O
d. 0 mol NO, 0 mol H_2, 4 mol NH_3, 4 mol H_2O

103. The circle shown here represents 1.0 liter of a solution with a solute concentration of 1 M:



Explain what you would add (the amount of solute or volume of solvent) to the solution above so as to obtain a solution represented by each of the following circles:

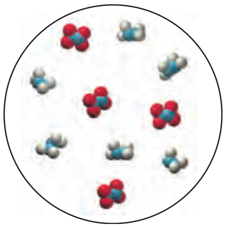


a. **b.** **c.**

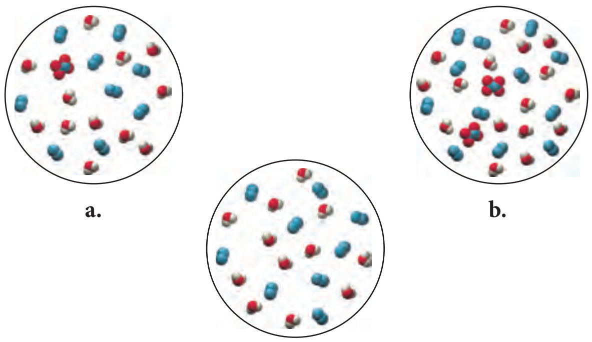
104. Consider the reaction.

$$2\,N_2H_4(g) + N_2O_4(g) \longrightarrow 3\,N_2(g) + 4\,H_2O(g)$$

Consider the following representation of an initial mixture of N_2H_4 and N_2O_4:



Which of the following diagrams best represents the reaction mixture after the reactants have reacted as completely as possible?

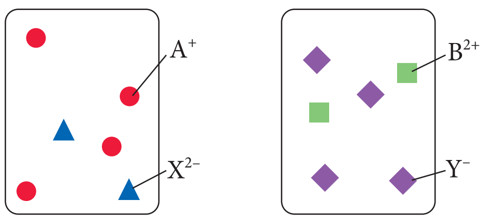


a. **b.** **c.**

105. Consider the generic ionic compounds with the formulas A_2X and BY_2 and the following solubility rules:

A_2X soluble; BY_2 soluble; AY insoluble; BX soluble

Let A^+ ions be represented by circles, B^{2+} ions be represented by squares, X^{2-} ions be represented by triangles, and Y^- ions be represented by diamonds. Solutions of the two compounds (A_2X and BY_2) can be represented as follows:



$A_2X(aq)$ $BY_2(aq)$

Draw a molecular-level representation showing the result of mixing the two solutions above and write an equation to represent the reaction.

Questions for Group Work

Discuss these questions with the group and record your consensus answer.

106. 16.05 g of methane (CH_4) gas is mixed with 96.00 g of oxygen (O_2) gas. The mixture is ignited. After a bright flash and a loud bang, some water droplets form on the inside of the reaction vessel.

a. Write the balanced chemical reaction for the combustion of methane.

b. Depict the process that occurred in the vessel using circles to represent atoms. Represent carbon with black circles, hydrogen with white circles, and oxygen with gray circles. Let one circle (or one molecule made of bonded circles) represent exactly one mole.

c. How many moles of water can you make? CO_2?

d. Will anything be left over? What? How much?

e. Identify the following: limiting reagent, excess reagent, and theoretical yield.

107. Write a detailed set of instructions for making two solutions: (1) 100.0 mL of 12 M NaOH from solid sodium hydroxide and (2) 1.00 L of 0.10 M NaOH from your first solution. You have in your lab: volumetric flasks marked to contain 100.0 mL and 1.000 L, a 25 mL graduated cylinder, and a balance.

108. Memorize the solubility rules. Without referring back to the rules, have each group member list two ionic compounds that are expected to be soluble and two that are expected to be insoluble. Include at least one exception. Check the work of the other members of your group.

109. Define and give an example of each class of reactions: precipitation, acid–base, gas evolution, redox (noncombustion), and combustion. Each group member can do one, and then present his or her reaction to the group.

110. Using group members to represent atoms, ions, or electrons, act out the reaction $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$. Which group member is oxidized? Which is reduced? What is the oxidizing agent? What is the reducing agent?

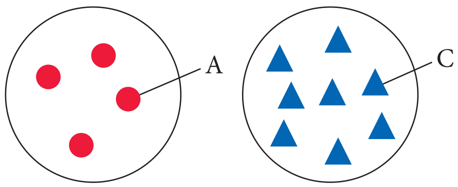
Answers to Conceptual Connections

Stoichiometry I

4.1. **(c)** Since each O_2 molecule reacts with 4 Na atoms, 12 Na atoms are required to react with 3 O_2 molecules.

Stoichiometry II

4.2.



Limiting Reactant and Theoretical Yield

4.3. **(c)** Nitrogen is the limiting reactant and there is enough nitrogen to make 4 NH_3 molecules. Hydrogen is in excess, and two hydrogen molecules remain after the reactants have reacted as completely as possible.

4.4. The limiting reactant is the 1 mol H_2O, which is completely consumed. The 1 mol of H_2O requires 3 mol of NO_2 to completely react; therefore 2 mol NO_2 remain after the reaction is complete.

Solutions

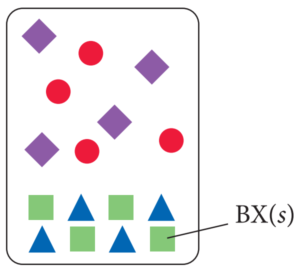
4.5. **(b)** The mass of a solution is equal to the mass of the solute plus the mass of the solvent. Although the solute seems to disappear, it really does not, and its mass becomes part of the mass of the solution, in accordance with the law of mass conservation.

Solution Dilution

4.6. **(c)** Since the volume has doubled, the concentration is halved, so the same volume should contain half as many solute molecules.

Precipitation Reactions

4.7.



$$AX(aq) + BY(aq) \longrightarrow BX(s) + AY(aq)$$

Oxidation and Reduction

4.8. **(d)** Since oxidation and reduction must occur together, an increase in the oxidation state of a reactant will always be accompanied by a decrease in the oxidation state of a reactant.