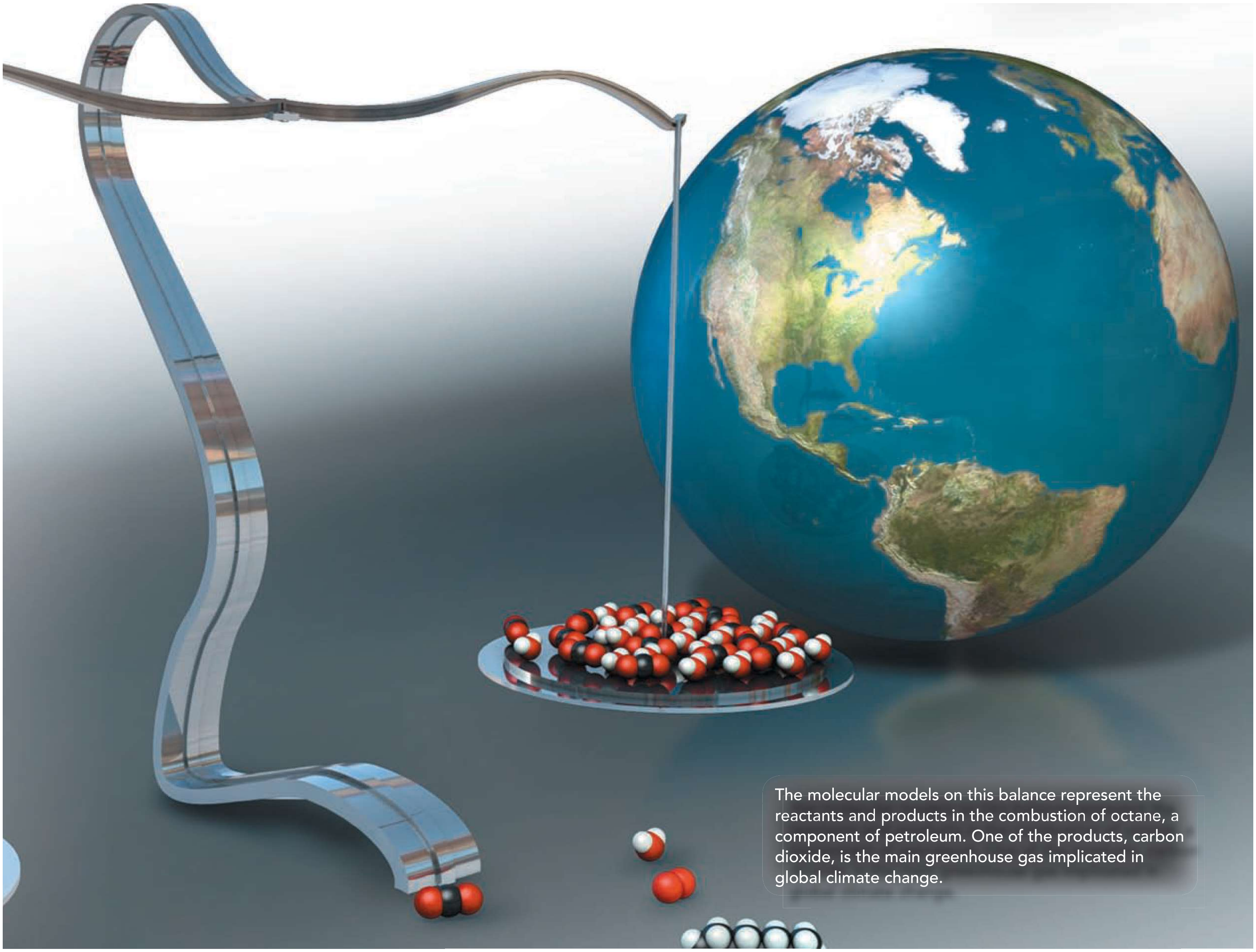
I feel sorry for people who don't know anything about chemistry. They are missing an important source of happiness.

—LINUS PAULING (1901–1994)

CHAPTER

4 Chemical Reactions and Chemical Quantities

In Chapter 3, we examined chemical compounds. We now turn to the process that can create and transform compounds: *chemical reactions*. We have seen that matter is composed of particles (atoms and molecules). When we mix certain types of particles with others, the electrons from one set of particles are attracted to the nuclei in the other set. If the conditions are right, a chemical reaction occurs and the particles are transformed. In this chapter, we learn to write chemical equations that represent these transformations. We also examine *chemical stoichiometry*—the numerical relationships between the amounts of reactants and products in chemical 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 global climate change.

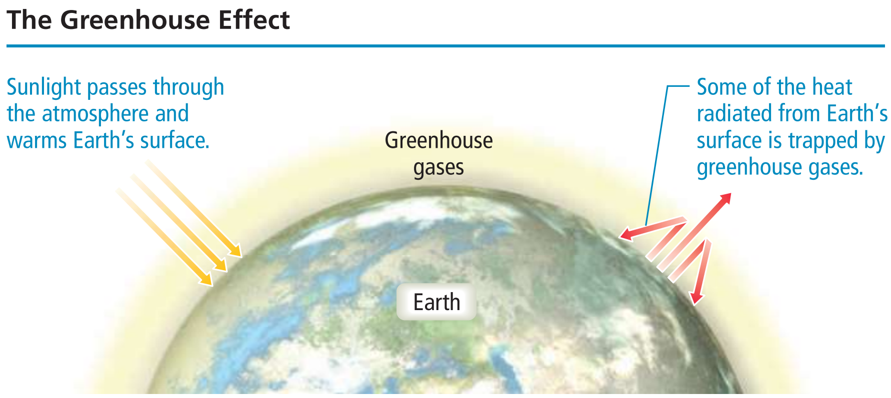
4.1 Climate Change and the Combustion of Fossil Fuels

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

Greenhouse gases in the atmosphere act as a one-way filter. They allow visible light to pass through and warm Earth's surface, but they prevent heat energy from radiating back out into space.

These gases act 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

▶ **FIGURE 4.1 The Greenhouse Effect**



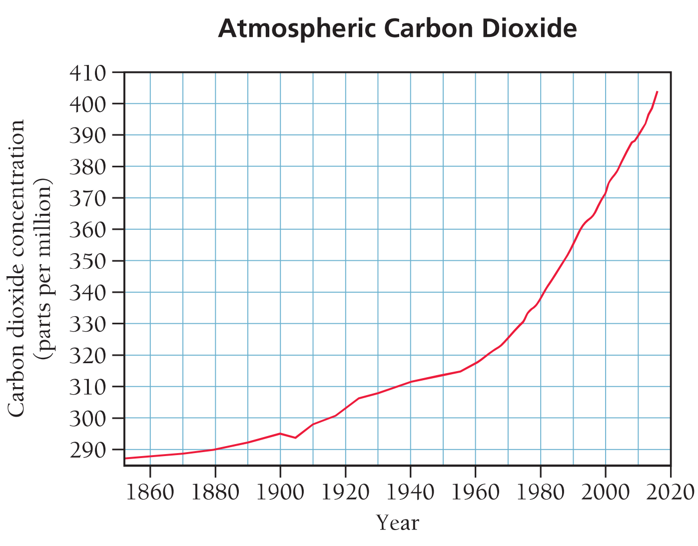
sunlight from escaping, as shown in Figure 4.1▲. The balance between incoming and outgoing energy from the sun determines Earth's average temperature.

If the greenhouse gases in the atmosphere were not present, more heat energy would escape, and Earth's average temperature would be about 60 °F colder than it currently is. 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.

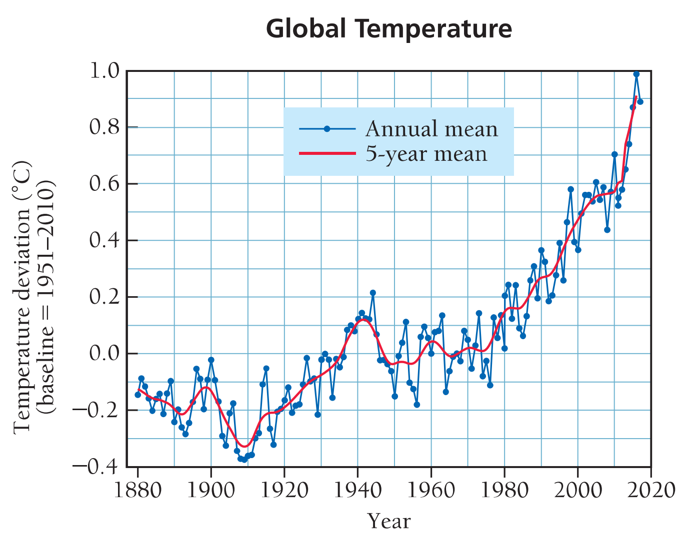
The extremely cold temperatures of Mars are as much a result of its lack of atmosphere as its greater distance from the sun than Earth. Conversely, Venus is an inferno partly because its thick atmosphere is rich in greenhouse gases.

In recent decades, scientists have become increasingly concerned because the quantity of atmospheric carbon dioxide (CO_2)—a significant greenhouse gas—is rising. More CO_2 enhances the atmosphere's ability to hold heat and leads 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 about 1.6 °F (0.9 °C) (as shown in Figure 4.3▼).

Most scientists believe that the primary cause of rising atmospheric CO_2 concentration is the burning of fossil fuels (natural gas, petroleum, and coal), which provides about 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 that the amount of carbon dioxide emitted into the atmosphere by natural sources, such as volcanic eruptions, far exceeds that from fossil fuel combustion. Which group is right? One way to judge the validity of the naysayers' argument is by calculating how much carbon dioxide is emitted by fossil fuel combustion and comparing that amount to the amount released by volcanic eruptions. In order to do this, we must understand the chemical equations that govern fossil fuel combustion and how those equations relate the amounts of reactants to the amounts of products.



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



▲ **FIGURE 4.3 Global Temperature** Average temperatures worldwide have risen by about 0.9 °C since 1880. Each point on the graph is the deviation from the 1951–2010 average temperature.

4.2 Writing and Balancing Chemical Equations

WATCH **NOW!**

KEY CONCEPT VIDEO 4.2
Writing and Balancing Chemical Equations

Combustion analysis (which we examined in Section 3.10) employs a **chemical reaction**, a process in which one or more substances are converted into one or more different ones. Compounds form and change through chemical reactions. For example, water is formed by the reaction of hydrogen with oxygen. A **combustion reaction** is a particular type of chemical reaction in which a substance combines with oxygen to form one or more oxygen-containing compounds. Combustion reactions also emit heat. The heat produced in a number of combustion reactions is critical to supplying our society's energy needs. For example, the heat from the combustion of gasoline expands the gaseous combustion products in a car engine's cylinders, which push the pistons and propel the car. We use the heat released by the combustion of *natural gas* to cook food and to heat our homes.

We represent a chemical reaction with a **chemical equation**. For example, we represent the combustion of natural gas with the equation:

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

The substances on the left side of the equation are the **reactants**, and the substances on the right side are the **products**. We often specify the states of each reactant or product in parentheses next to the formula as follows:

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

The (g) indicates that these substances are gases in the reaction. Table 4.1 summarizes the common states of reactants and products and their symbols used in chemical equations.

TABLE 4.1 ■ States of Reactants and Products in Chemical Equations

Abbreviation	State
(g)	Gas
(l)	Liquid
(s)	Solid
(aq)	Aqueous (water solution)

The equation just presented for the combustion of natural gas is not complete, however. If we look closely, we can immediately spot a problem.

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

2 O atoms

2 O atoms + 1 O atom = **3 O atoms**

The left side of the equation has two oxygen atoms, while the right side has three. The reaction as written, therefore, violates the law of conservation of mass because an oxygen atom formed out of nothing. Notice also that the left side has four hydrogen atoms, while the right side has only two.

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

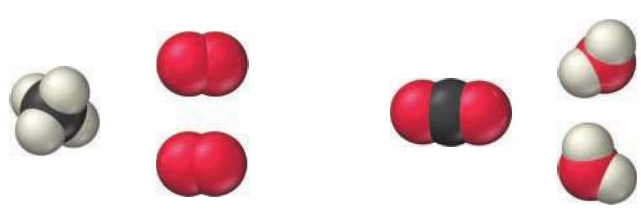
4 H atoms **2 H atoms**

Two hydrogen atoms have vanished, again violating mass conservation. To correct these problems—that is, to write an equation that more closely represents *what actually happens*—we must **balance** the equation. We need to change the coefficients (the numbers *in front of* the chemical formulas), not the subscripts (the numbers within the chemical formulas), to ensure that the number of each type of atom on the left side of the equation is equal to the number on the right side. New atoms do not form during a reaction, nor do atoms vanish—matter must be conserved.

We cannot change the subscripts when balancing a chemical equation because changing the subscripts changes the substance itself, while changing the coefficients changes the number of molecules of the substance. For example, 2 H_2O is simply two water molecules, but H_2O_2 is hydrogen peroxide, a drastically different compound.

When we add coefficients to the reactants and products to balance an equation, we change the number of molecules in the equation but not the *kind of* molecules. To balance the equation for the combustion of methane, we put the coefficient 2 before O_2 in the reactants, and the coefficient 2 before H_2O in the products.

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



The equation is now balanced because the numbers of each type of atom on either side of the equation are equal. The balanced equation tells us that one CH_4 molecule reacts with two O_2 molecules to form one CO_2 molecule and two H_2O molecules. We verify that the equation is balanced by summing the number of each type of atom on each side of the equation.

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

Reactants	Products
1 C atom ($1 \times \underline{C}H_4$)	1 C atom ($1 \times \underline{C}O_2$)
4 H atoms ($1 \times C\underline{H_4}$)	4 H atoms ($2 \times \underline{H_2}O$)
4 O atoms ($2 \times \underline{O_2}$)	4 O atoms ($1 \times C\underline{O_2} + 2 \times H_2\underline{O}$)

The number of each type of atom on both sides of the equation is now equal—the equation is balanced.

ANSWER **NOW!**

4.1 Cc Conceptual Connection

COUNTING ATOMS IN A CHEMICAL EQUATION

How many oxygen atoms are on the right-hand side of the following chemical equation?

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

(a) 4 **(b)** 5 **(c)** 6 **(d)** 14

We can balance many chemical equations simply by trial and error. However, some guidelines are useful. For example, balancing the atoms in the most complex substances first and the atoms in the simplest substances (such as pure elements) last often makes the process shorter. The following examples of how to balance chemical equations are presented in a two- or three-column format. The general procedure is shown on the left, with the application of the procedure on the right. This procedure is meant only as a flexible guide, not a rigid set of steps.

WATCH **NOW!** INTERACTIVE WORKED EXAMPLE 4.2

HOW TO: *Balance Chemical Equations*	**EXAMPLE 4.1** Balancing Chemical Equations	**EXAMPLE 4.2** Balancing Chemical Equations
	Write a balanced equation for the reaction between solid cobalt(III) oxide and solid carbon to produce solid cobalt and carbon dioxide gas.	Write a balanced equation for the combustion of gaseous butane (C_4H_{10}), a fuel used in portable stoves and grills, in which it combines with gaseous oxygen to form gaseous carbon dioxide and gaseous water.
1. Write a skeletal equation by writing chemical formulas for each of the reactants and products. Review Sections 3.5 and 3.6 for nomenclature rules. (If a skeletal equation is provided, go to step 2.)	$Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$	$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$

2. Balance atoms that occur in more complex substances first. Always balance atoms in compounds before atoms in pure elements.

Begin with O:

$Co_2O_3(s) + C(s) \longrightarrow Co(s) + CO_2(g)$

3 O atoms ⟶ 2 O atoms

To balance O, put a 2 before $Co_2O_3(s)$ and a 3 before $CO_2(g)$.

$2\ Co_2O_3(s) + C(s) \longrightarrow Co(s) + 3\ CO_2(g)$

6 O atoms ⟶ 6 O atoms

Begin with C:

$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$

4 C atoms ⟶ 1 C atom

To balance C, put a 4 before $CO_2(g)$.

$C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + H_2O(g)$

4 C atoms ⟶ 4 C atoms

Balance H:

$C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + H_2O(g)$

10 H atoms ⟶ 2 H atoms

To balance H, put a 5 before $H_2O(g)$:

$C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$

10 H atoms ⟶ 10 H atoms

3. Balance atoms that occur as free elements on either side of the equation last. Balance free elements by adjusting their coefficients.

Balance Co:

$2\ Co_2O_3(s) + C(s) \longrightarrow Co(s) + 3\ CO_2(g)$

4 Co atoms ⟶ 1 Co atom

To balance Co, put a 4 before $Co(s)$.

$2\ Co_2O_3(s) + C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$

4 Co atoms ⟶ 4 Co atoms

Balance C:

$2\ Co_2O_3(s) + C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$

1 C atom ⟶ 3 C atoms

To balance C, put a 3 before $C(s)$.

$2\ Co_2O_3(s) + 3\ C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$

Balance O:

$C_4H_{10}(g) + O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$

2 O atoms ⟶ 8 O + 5 O = 13 O atoms

To balance O, put a 13/2 before $O_2(g)$.

$C_4H_{10}(g) + 13/2\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)$

13 O atoms ⟶ 13 O atoms

4. If the balanced equation contains coefficient fractions, clear these by multiplying each of the coefficients in the entire equation by the denominator of the fraction.

This step is not necessary in this example. Proceed to step 5.

$[C_4H_{10}(g) + 13/2\ O_2(g) \longrightarrow 4\ CO_2(g) + 5\ H_2O(g)] \times 2$

$2\ C_4H_{10}(g) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g)$

5. Check to make certain the equation is balanced by summing the total number of each type of atom on both sides of the equation.

$2\ Co_2O_3(s) + 3\ C(s) \longrightarrow 4\ Co(s) + 3\ CO_2(g)$

Left	Right
4 Co atoms	4 Co atoms
6 O atoms	6 O atoms
3 C atoms	3 C atoms

The equation is balanced.

$2\ C_4H_{10}(g) + 13\ O_2(g) \longrightarrow 8\ CO_2(g) + 10\ H_2O(g)$

Left	Right
8 C atoms	8 C atoms
20 H atoms	20 H atoms
26 O atoms	26 O atoms

The equation is balanced.

FOR PRACTICE 4.1

Write a balanced equation for the reaction between solid silicon dioxide and solid carbon to produce solid silicon carbide and carbon monoxide gas.

FOR PRACTICE 4.2

Write a balanced equation for the combustion of gaseous ethane (C_2H_6), a minority component of natural gas, in which it combines with gaseous oxygen to form gaseous carbon dioxide and gaseous water.

ANSWER **NOW!**

4.2 Cc Conceptual Connection

BALANCED CHEMICAL EQUATIONS

Which quantity or quantities must always be the same on both sides of a chemical equation?

(a) the number of atoms of each kind
(b) the number of molecules of each kind
(c) the number of moles of each kind of molecule

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 4.3

EXAMPLE 4.3 Balancing Chemical Equations Containing Polyatomic Ions

Write a balanced equation for the reaction between aqueous strontium chloride and aqueous lithium phosphate to form solid strontium phosphate and aqueous lithium chloride.

SOLUTION

1. Write a skeletal equation by writing chemical formulas for each of the reactants and products. Review Sections 3.5 and 3.6 for naming rules. (If a skeletal equation is provided, go to step 2.)	$SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$
2. Balance metal ions (cations) first. If a polyatomic cation exists on both sides of the equation, balance it as a unit.	**Begin with Sr^{2+}:** $SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 1 Sr^{2+} ion $\longrightarrow$ 3 Sr^{2+} ions To balance Sr^{2+}, put a 3 before $SrCl_2(aq)$. $3\,SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 3 Sr^{2+} ions $\longrightarrow$ 3 Sr^{2+} ions **Balance Li^+:** $3\,SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + LiCl(aq)$ 3 Li^+ ions $\longrightarrow$ 1 Li^+ ion To balance Li^+, put a 3 before $LiCl(aq)$. $3\,SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\,LiCl(aq)$ 3 Li^+ ions $\longrightarrow$ 3 Li^+ ions
3. Balance nonmetal ions (anions) second. If a polyatomic anion exists on both sides of the equation, balance it as a unit.	**Balance PO_4^{3-}:** $3\,SrCl_2(aq) + Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\,LiCl(aq)$ 1 PO_4^{3-} ion $\longrightarrow$ 2 PO_4^{3-} ions To balance PO_4^{3-}, put a 2 before $Li_3PO_4(aq)$. $3\,SrCl_2(aq) + 2\,Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\,LiCl(aq)$ 2 PO_4^{3-} ions $\longrightarrow$ 2 PO_4^{3-} ions **Balance Cl^-:** $3\,SrCl_2(aq) + 2\,Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 3\,LiCl(aq)$ 6 Cl^- ions $\longrightarrow$ 3 Cl^- ions To balance Cl^-, replace the 3 before $LiCl(aq)$ with a 6. This also corrects the balance for Li^+, which was thrown off in the previous step. $3\,SrCl_2(aq) + 2\,Li_3PO_4(aq) \longrightarrow Sr_3(PO_4)_2(s) + 6\,LiCl(aq)$ 6 Cl^- ions $\longrightarrow$ 6 Cl^- ions

4. Check to make certain the equation is balanced by summing the total number of each type of ion on both sides of the equation.

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

Left	Right
3 Sr^{2+} ions	3 Sr^{2+} ions
6 Li^{+} ions	6 Li^{+} ions
2 PO_4^{3-} ions	2 PO_4^{3-} ions
6 Cl^{-} ions	6 Cl^{-} ions

The equation is balanced.

FOR PRACTICE 4.3 Write a balanced equation for the reaction between aqueous lead(II) nitrate and aqueous potassium chloride to form solid lead(II) chloride and aqueous potassium nitrate.

4.3 Reaction Stoichiometry: How Much Carbon Dioxide?

WATCH **NOW!**

KEY CONCEPT VIDEO 4.3
Reaction Stoichiometry

The balanced chemical equations for fossil fuel combustion reactions provide the exact relationships between the amount of fossil fuel burned and the amount of carbon dioxide emitted. 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 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 equation 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 relationships between chemical amounts in a balanced chemical equation are called 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 react. Stoichiometry also allows us to determine the amounts 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 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 from 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 contains the numerical relationships between the pizza ingredients. It says that if we have two cups of cheese—and enough of everything else—we can make one 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 six cups of cheese? Assuming that we have enough of everything else, we can use the 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 three pizzas. The pizza recipe contains numerical ratios between other ingredients as well, including the crust and the tomato sauce:

1 crust : 1 pizza

5 ounces tomato sauce : 1 pizza

Making Molecules: Mole-to-Mole Conversions

In a balanced chemical equation, we have 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:

$$2 \text{ mol } C_8H_{18} : 16 \text{ mol } CO_2$$

We can use this ratio to determine how many moles of CO_2 form when a given number of moles of C_8H_{18} burns. Suppose that we burn 22.0 moles of C_8H_{18}; how many moles of CO_2 form? We use the ratio from the balanced chemical equation in the same way that we used the ratio from the pizza recipe. The ratio acts as a conversion factor between the amount in moles of the reactant (C_8H_{18}) and 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 22.0 moles of C_8H_{18} adds 176 moles of CO_2 to the atmosphere.

ANSWER **NOW!**

4.3 Cc Conceptual Connection

STOICHIOMETRY I Use the balanced equation for the combustion of octane to determine how many moles of H_2O are produced by the combustion of 22.0 moles of C_8H_{18}.

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

(a) 18 moles H_2O **(b)** 22 moles H_2O
(c) 176 moles H_2O **(d)** 198 moles H_2O

Making Molecules: Mass-to-Mass Conversions

According to the U.S. Department of Energy, the world burned 3.6×10^{10} barrels of petroleum in 2017, the equivalent of approximately 4.0×10^{15} g of gasoline. We can estimate the mass of CO_2 emitted into the atmosphere from burning this much gasoline using the combustion of 4.0×10^{15} g octane as the representative reaction. This calculation is similar to the one we just did, except that we are now 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 takes the form:

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

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 4.0×10^{15} g of octane, 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 mol C_8H_{18} : 16 mol CO_2 (from the chemical equation)

molar mass C_8H_{18} = 114.22 g/mol

molar mass CO_2 = 44.01 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:

$$4.0 \times 10^{15}\ \text{g}\ \cancel{C_8H_{18}} \times \frac{1\ \cancel{\text{mol}\ C_8H_{18}}}{114.22\ \cancel{\text{g}\ C_8H_{18}}} \times \frac{16\ \cancel{\text{mol}\ CO_2}}{2\ \cancel{\text{mol}\ C_8H_{18}}} \times \frac{44.01\ \text{g}\ CO_2}{1\ \cancel{\text{mol}\ CO_2}} = 1.2 \times 10^{16}\ \text{g}\ CO_2$$

The world's petroleum combustion produces 1.2×10^{16} g CO_2 (1.2×10^{13} kg) per year. In comparison, volcanoes produce about 2×10^{11} kg CO_2 per year.* In other words, volcanoes emit only $\frac{2 \times 10^{11}\ \text{kg}}{1.2 \times 10^{13}\ \text{kg}} \times 100\% = 1.7\%$ as much CO_2 per year as petroleum combustion. The argument that volcanoes emit more carbon dioxide than fossil fuel combustion is clearly mistaken. Examples 4.4 and 4.5 provide additional practice with stoichiometric calculations.

The percentage of CO_2 emitted by volcanoes relative to all fossil fuels is even less than 1.7% because the combustion of coal and natural gas also emits CO_2.

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 4.4

EXAMPLE 4.4 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 that a particular plant consumes 37.8 g of 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 provides 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 → amount A (in moles) → amount B (in moles) → mass B. From the chemical equation, deduce the relationship between moles of carbon dioxide and moles of glucose. Use the molar masses to convert between grams and moles.

CONCEPTUAL PLAN

$$\text{g}\ CO_2 \xrightarrow{\frac{1\ \text{mol}\ CO_2}{44.01\ \text{g}\ CO_2}} \text{mol}\ CO_2 \xrightarrow{\frac{1\ \text{mol}\ C_6H_{12}O_6}{6\ \text{mol}\ CO_2}} \text{mol}\ C_6H_{12}O_6 \xrightarrow{\frac{180.2\ \text{g}\ C_6H_{12}O_6}{1\ \text{mol}\ C_6H_{12}O_6}} \text{g}\ C_6H_{12}O_6$$

RELATIONSHIPS USED

molar mass CO_2 = 44.01 g/mol

6 mol CO_2 : 1 mol $C_6H_{12}O_6$

molar mass $C_6H_{12}O_6$ = 180.2 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

$$37.8\ \text{g}\ \cancel{CO_2} \times \frac{1\ \cancel{\text{mol}\ CO_2}}{44.01\ \cancel{\text{g}\ CO_2}} \times \frac{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}{6\ \cancel{\text{mol}\ CO_2}} \times \frac{180.2\ \text{g}\ C_6H_{12}O_6}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}} = 25.8\ \text{g}\ C_6H_{12}O_6$$

—Continued on the next page

*Gerlach, T. M., Present-day CO_2 emissions from volcanoes. *Eos, Transactions, American Geophysical Union* 72(1991): 249, 254–255.

Continued—

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

$$Mg(OH)_2(s) + 2\,HCl(aq) \longrightarrow 2\,H_2O(l) + MgCl_2(aq)$$

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

EXAMPLE 4.5 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 by a 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?

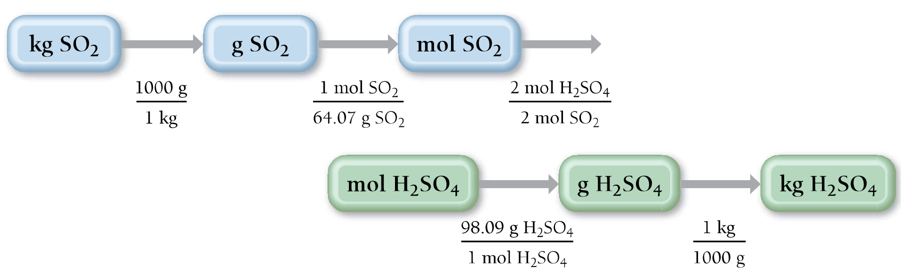
SORT The problem gives the mass of sulfur dioxide and asks you to find the mass of sulfuric acid.

GIVEN: 25 kg SO_2

FIND: kg H_2SO_4

STRATEGIZE The conceptual plan follows the standard format of mass → amount (in moles) → amount (in moles) → mass. Since the original quantity of SO_2 is given in kilograms, 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 kilograms, convert to kilograms at the end.

CONCEPTUAL PLAN



RELATIONSHIPS USED

1 kg = 1000 g

molar mass SO_2 = 64.07 g/mol

2 mol SO_2 : 2 mol H_2SO_4

molar mass H_2SO_4 = 98.09 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

$$25\,\cancel{kg\,SO_2} \times \frac{1000\,\cancel{g}}{1\,\cancel{kg}} \times \frac{1\,\cancel{mol\,SO_2}}{64.07\,\cancel{g\,SO_2}} \times \frac{2\,\cancel{mol\,H_2SO_4}}{2\,\cancel{mol\,SO_2}} \times \frac{98.09\,\cancel{g}\,H_2SO_4}{1\,\cancel{mol\,H_2SO_4}} \times \frac{1\,kg}{1000\,\cancel{g}} = 38\,kg\,H_2SO_4$$

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 4.5 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:

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

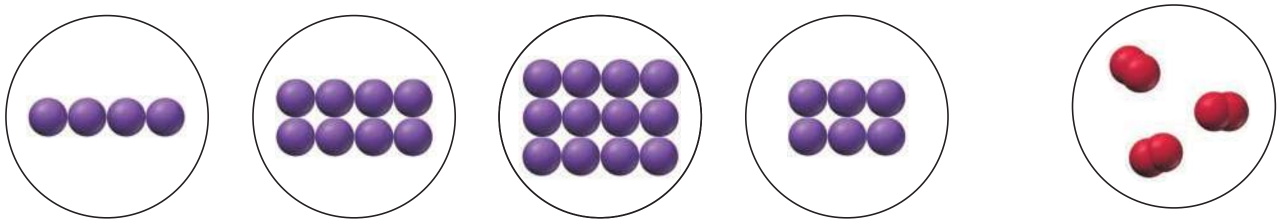
The generation of the electricity used by a medium-sized home produces about 16 kg of NO_2 per year. Assuming that there is adequate O_2 and H_2O, what mass of HNO_3, in kg, can form from this amount of NO_2 pollutant?

STOICHIOMETRY II Under certain conditions, sodium reacts with oxygen to form sodium oxide according to the reaction:

$$4\ \mathrm{Na}(s) + \mathrm{O_2}(g) \longrightarrow 2\ \mathrm{Na_2O}(s)$$

A flask contains the amount of oxygen represented by the diagram shown at far right.

Which of the following images best represents the amount of sodium required to completely react with all of the oxygen in the flask according to the equation?

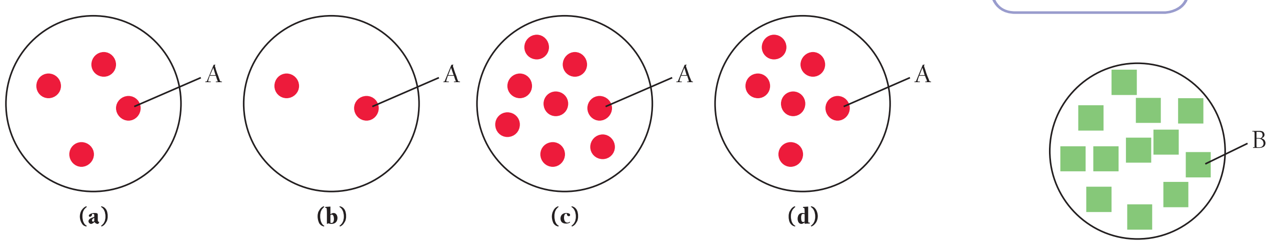


(a) **(b)** **(c)** **(d)**

4.4 Cc Conceptual Connection

ANSWER **NOW!**

STOICHIOMETRY III Consider the generic chemical equation A + 3B → 2C. Let circles represent molecules of A and squares represent molecules of B. The diagram shown at the far right represents the amount of B available for reaction. Which diagram in the answer options accurately represents the amount of A necessary to completely react with B?



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

4.5 Cc Conceptual Connection

ANSWER **NOW!**

4.4 Stoichiometric Relationships: Limiting Reactant, Theoretical Yield, Percent Yield, and Reactant in Excess

WATCH **NOW!**

KEY CONCEPT VIDEO 4.4
Limiting Reactant, Theoretical Yield, and Percent Yield

Let's return to our pizza analogy to understand three more important concepts in reaction stoichiometry: *limiting reactant, theoretical yield*, and *percent yield*. Recall our pizza recipe from Section 4.3:

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

Suppose that we have four crusts, ten 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:

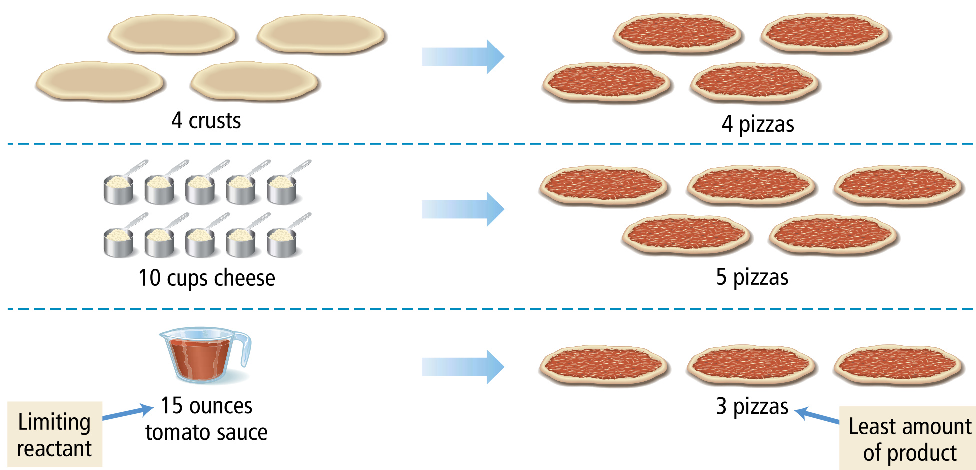
$$10\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$$

We have enough tomato sauce to make:

$$\underbrace{15\ \cancel{\text{ounces tomato sauce}}}_{\text{Limiting reactant}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = \underbrace{3\ \text{pizzas}}_{\text{Smallest number of pizzas}}$$

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

We have enough crusts for four pizzas and enough cheese for five pizzas, but enough tomato sauce for only three pizzas. Consequently, unless we get more ingredients, we can make only three 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 the reactant that makes *the least amount of product*.



The ingredient that makes the least amount of pizza determines how many pizzas you can make.

The 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, three pizzas would be the **theoretical yield**, the maximum 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 three pizzas, we end up with only two. If this were a chemical reaction, the two 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 a small amount of product is usually lost to other reactions or does not form during a reaction.) Finally, our **percent yield**, the percentage of the theoretical yield that was actually attained, is calculated as the ratio of the actual yield to the theoretical yield:

$$\%\ \text{yield} = \frac{\overbrace{2\ \cancel{\text{pizzas}}}^{\text{Actual yield}}}{\underbrace{3\ \cancel{\text{pizzas}}}_{\text{Theoretical yield}}} \times 100\% = 67\%$$

Since one of our pizzas burned, our percent yield for pizzas is 67%.

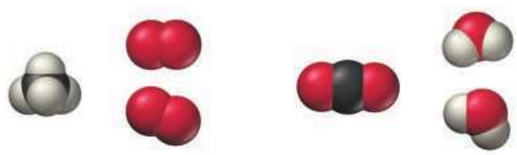
Summarizing Limiting Reactant and Yield

- **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 $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$.

Calculating Limiting Reactant, Theoretical Yield, and Percent Yield

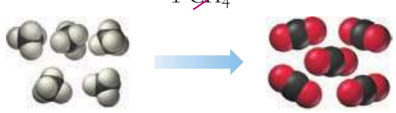
We can apply these concepts to a chemical reaction. Recall from Section 4.2 our balanced equation for the combustion of methane:

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



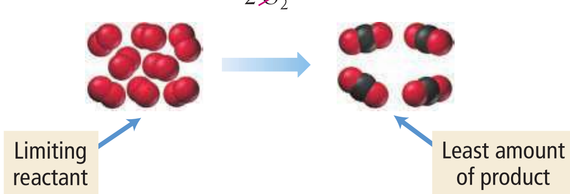
If we start out with five CH_4 molecules and eight O_2 molecules, what is our limiting reactant? What is our theoretical yield of carbon dioxide molecules? First, we calculate the number of CO_2 molecules that can be made from five CH_4 molecules:

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



We then calculate the number of CO_2 molecules that can be made from eight 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 five CO_2 molecules and enough O_2 to make four CO_2 molecules; therefore, O_2 is the limiting reactant, and four CO_2 molecules is the theoretical yield. The CH_4 is in excess.

An alternative way to calculate the limiting reactant (which we describe here but do not use again in this book) is to pick any reactant and determine how much of the *other reactant* is necessary to completely react with it. For the reaction we just examined, we have five CH_4 molecules and eight O_2 molecules. Let's pick the five CH_4 molecules and determine how many O_2 molecules are necessary to completely react with them:

$$5\,\cancel{CH_4} \times \frac{2\,O_2}{1\,\cancel{CH_4}} = 10\,O_2$$

Since we need ten O_2 molecules to completely react with the five CH_4 molecules, and since we have only eight O_2 molecules, we know that the O_2 is the limiting reactant. The same method can be applied by comparing the amounts of reactants in moles.

ANSWER NOW!

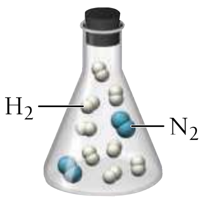
4.6 Cc Conceptual Connection

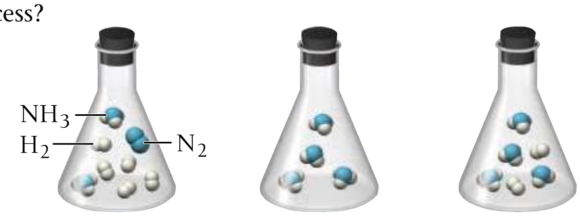
LIMITING REACTANT AND THEORETICAL YIELD Nitrogen and hydrogen gas react to form ammonia according to the reaction:

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

A flask contains a mixture of reactants represented by the image shown at the left.

Which of the following images best represents the mixture in the flask after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?





(a) (b) (c)

Calculating 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 the reaction:

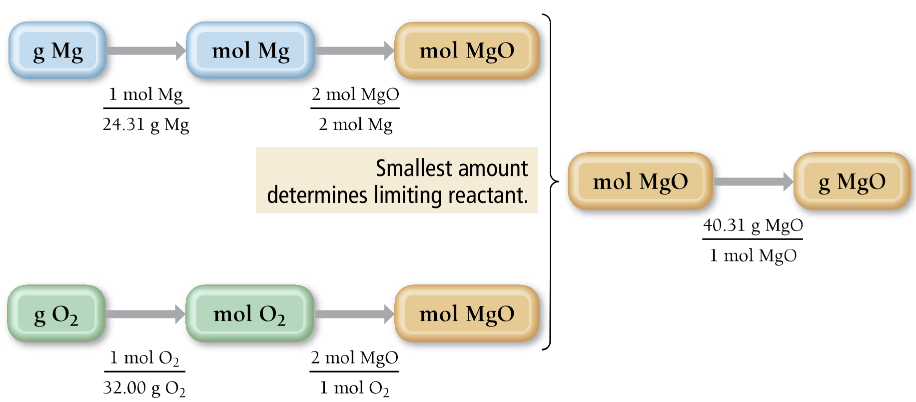
$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

A reaction mixture contains 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 can find the limiting reactant by calculating how much product can be made from each reactant. However, we are given the initial quantities in grams, and stoichiometric relationships are between moles, so 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:

$$42.5\ \cancel{\text{g Mg}} \times \frac{1\ \cancel{\text{mol Mg}}}{24.31\ \cancel{\text{g Mg}}} \times \frac{2\ \text{mol MgO}}{2\ \cancel{\text{mol Mg}}} = 1.7483\ \text{mol MgO}$$

Limiting reactant

Least amount of product

$$33.8\ \cancel{\text{g O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g O}_2}} \times \frac{2\ \text{mol MgO}}{1\ \cancel{\text{mol O}_2}} = 2.1125\ \text{mol MgO}$$

$$1.7483\ \cancel{\text{mol MgO}} \times \frac{40.31\ \text{g MgO}}{1\ \cancel{\text{mol MgO}}} = 70.5\ \text{g MgO}$$

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 70.5 g of MgO, the mass of product possible based on the limiting reactant.

Suppose that after the synthesis, the actual yield of MgO is 55.9 g. What is the percent yield? We calculate the percent yield as follows:

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

WATCH **NOW!**

INTERACTIVE WORKED EXAMPLE 4.6

EXAMPLE 4.6 Limiting Reactant and Theoretical Yield

Ammonia, NH_3, can be synthesized by the reaction:

$$2\,\text{NO}(g) + 5\,\text{H}_2(g) \longrightarrow 2\,\text{NH}_3(g) + 2\,\text{H}_2\text{O}(g)$$

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 $NH_3(g)$

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 (from the chemical equation) to convert between moles of reactant and moles of product. Remember that 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

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

$$\text{g H}_2 \xrightarrow{\frac{1\ \text{mol H}_2}{2.02\ \text{g H}_2}} \text{mol H}_2 \xrightarrow{\frac{2\ \text{mol NH}_3}{5\ \text{mol H}_2}} \text{mol NH}_3$$

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 $H_2 = 2.02$ 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

—*Continued on the next page*

Continued—

SOLVE Beginning with the given mass 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 grams—this is the theoretical yield.

SOLUTION

$$\underbrace{86.3\ \cancel{\text{g NO}}}_{\text{Limiting reactant}} \times \frac{1\ \cancel{\text{mol NO}}}{30.01\ \cancel{\text{g NO}}} \times \frac{2\ \text{mol NH}_3}{2\ \cancel{\text{mol NO}}} = \underbrace{2.87\underline{5}7\ \text{mol NH}_3}_{\text{Least amount of product}}$$

$$25.6\ \cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.02\ \cancel{\text{g H}_2}} \times \frac{2\ \text{mol NH}_3}{5\ \cancel{\text{mol H}_2}} = 5.0\underline{6}93\ \text{mol NH}_3$$

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

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

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 three hydrogen atoms per nitrogen atom. Since three hydrogen atoms have less mass than one oxygen atom, it is reasonable that the mass of NH_3 obtained is less than the mass of NO.

FOR PRACTICE 4.6 Ammonia can also be synthesized by the reaction:

$$3\ \text{H}_2(g) + \text{N}_2(g) \longrightarrow 2\ \text{NH}_3(g)$$

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

EXAMPLE 4.7 Limiting Reactant and Theoretical Yield

We can obtain titanium metal from its oxide according to the following balanced equation:

$$\text{TiO}_2(s) + 2\ \text{C}(s) \longrightarrow \text{Ti}(s) + 2\ \text{CO}(g)$$

When 28.6 kg of C reacts with 88.2 kg of TiO_2, 42.8 kg of Ti is produced. Find 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. Remember that the reactant that makes the *least amount of product* is the limiting reactant.

CONCEPTUAL PLAN

$$\text{kg C} \xrightarrow{\frac{1000\ \text{g}}{1\ \text{kg}}} \text{g C} \xrightarrow{\frac{1\ \text{mol C}}{12.01\ \text{g C}}} \text{mol C} \xrightarrow{\frac{1\ \text{mol Ti}}{2\ \text{mol C}}} \text{mol Ti}$$

$$\text{kg TiO}_2 \xrightarrow{\frac{1000\ \text{g}}{1\ \text{kg}}} \text{g TiO}_2 \xrightarrow{\frac{1\ \text{mol TiO}_2}{79.87\ \text{g TiO}_2}} \text{mol TiO}_2 \xrightarrow{\frac{1\ \text{mol Ti}}{1\ \text{mol TiO}_2}} \text{mol Ti}$$

Smallest amount determines limiting reactant.

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

$$\text{mol Ti} \xrightarrow{\frac{47.87\ \text{g Ti}}{\text{mol Ti}}} \text{g Ti} \xrightarrow{\frac{1\ \text{kg}}{1000\ \text{g}}} \text{kg Ti}$$

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

$$28.6\ \cancel{\text{kg C}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol C}}}{12.01\ \cancel{\text{g C}}} \times \frac{1\ \text{mol Ti}}{2\ \cancel{\text{mol C}}} = 1.1\underline{9}07 \times 10^3\ \text{mol Ti}$$

Limiting reactant

Least amount of product

$$\boxed{88.2\ \cancel{\text{kg TiO}_2}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \cancel{\text{mol TiO}_2}}{79.87\ \cancel{\text{g TiO}_2}} \times \frac{1\ \text{mol Ti}}{1\ \cancel{\text{mol TiO}_2}} = \boxed{1.1\underline{0}43 \times 10^3\ \text{mol Ti}}$$

$$1.10\underline{4}3 \times 10^3\ \cancel{\text{mol Ti}} \times \frac{47.87\ \cancel{\text{g Ti}}}{1\ \cancel{\text{mol Ti}}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 52.9\ \text{kg Ti}$$

Calculate the percent yield by dividing the actual yield (42.8 kg Ti) by the theoretical yield.

Since TiO_2 makes the least amount of product, it is the limiting reactant, and 52.9 kg Ti is the theoretical yield.

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{42.8\ \text{kg}}{52.9\ \text{kg}} \times 100\% = 80.9\%$$

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 4.7 Mining companies use this reaction 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

4.7 Cc Conceptual Connection

ANSWER **NOW!**

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

(a) 1 mol NO_2 **(b)** 1 mol H_2O **(c)** 2 mol NO_2 **(d)** 2 mol H_2O

4.5 Three Examples of Chemical Reactions: Combustion, Alkali Metals, and Halogens

In this section, we examine three types of reactions. The first is combustion reactions, which we encountered in Section 4.2. The second is the reactions of the alkali metals. As we discussed in Section 2.7, alkali metals are among the most active metals. Alkali metal reactions are good examples of the types of reactions that many metals undergo. The third type of reactions involves the halogens. Halogens are among the most active nonmetals.

Combustion Reactions

A *combustion reaction* involves 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 you saw earlier in this chapter, natural gas (CH_4) reacts with oxygen to form carbon dioxide and water:

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

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:

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

We can write chemical equations for most combustion reactions by noticing the pattern of reactivity. Any carbon in a combustion reaction reacts with oxygen to produce carbon dioxide, and any hydrogen reacts with oxygen to form water.

EXAMPLE 4.8 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 using the guidelines from Section 4.2.	$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$

FOR PRACTICE 4.8 Write a balanced equation for the complete combustion of liquid C_2H_5SH. *Hint:* The sulfur in this compound reacts to form SO_2.

Alkali Metal Reactions

The reactions of the alkali metals (group 1A in the periodic table) with nonmetals are vigorous. For example, the alkali metals (M) react with halogens (X) according to the reaction:

$$2\,M + X_2 \longrightarrow 2\,MX$$

The reaction of sodium and chlorine to form sodium chloride is typical:

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

This reaction emits heat and sparks as it occurs (Figure 4.4◀). Each successive alkali metal reacts even more vigorously with chlorine.

▲ **FIGURE 4.4 Reaction of Sodium and Chlorine to Form Sodium Chloride**

The alkali metals also react with water to form the dissolved alkali metal ion, the hydroxide ion, and hydrogen gas:

$$2\,M(s) + 2\,H_2O(l) \longrightarrow 2\,M^+(aq) + 2\,OH^-(aq) + H_2(g)$$

The reaction is highly exothermic and can be explosive because the heat from the reaction can ignite the hydrogen gas. The reaction becomes more explosive as we move down the column from one metal to the next, as shown in Figure 4.5▶.*

Halogen Reactions

The halogens (group 7A) are among the most active nonmetals in the periodic table. The halogens all react with many metals to form *metal halides* according to the equation:

$$2\,M + n\,X_2 \longrightarrow 2\,MX_n$$

*The rate of the alkali metal reaction with water, and therefore its vigor, is enhanced by the successively lower melting points of the alkali metals as we move down the column. The low melting points of the heavier metals allow the emitted heat to actually melt the metal, increasing the reaction rate.

Reactions of the Alkali Metals with Water

◀ **FIGURE 4.5 Reactions of the Alkali Metals with Water** The reactions become progressively more vigorous as we move down the group.

where M is the metal, X is the halogen, and MX_n is the metal halide. For example, chlorine reacts with iron according to the equation:

$$2\,Fe(s) + 3\,Cl_2(g) \longrightarrow 2\,FeCl_3(s)$$

Since metals tend to lose electrons and the halogens tend to gain them, the metal halides—like all compounds that form between metals and nonmetals—contain ionic bonds.

The halogens also react with hydrogen to form *hydrogen halides* according to the equation:

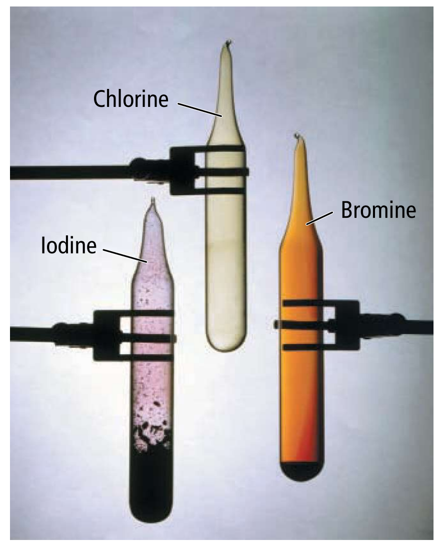
$$H_2(g) + X_2 \longrightarrow 2\,HX(g)$$

The hydrogen halides—like all compounds that form between two nonmetals—contain covalent bonds. All of the hydrogen halides form acids when combined with water.

The halogens also react with each other to form *interhalogen compounds*. For example, bromine reacts with fluorine according to the equation:

$$Br_2(l) + F_2(g) \longrightarrow 2\,BrF(g)$$

Again, like all compounds that form between two nonmetals, the interhalogen compounds contain covalent bonds.



▲ Three Halogens

EXAMPLE 4.9 Alkali Metal and Halogen Reactions

Write a balanced chemical equation for each reaction.

(a) the reaction between potassium metal and bromine gas

(b) the reaction between rubidium metal and liquid water

(c) the reaction between gaseous chlorine and solid iodine

SOLUTION

(a) Alkali metals react with halogens to form metal halides. Write the formulas for the reactants and the metal halide product (making sure to write the correct ionic chemical formula for the metal halide, as outlined in Section 3.5), and then balance the equation.	$2\,K(s) + Br_2(g) \longrightarrow 2\,KBr(s)$
(b) Alkali metals react with water to form the dissolved metal ion, the hydroxide ion, and hydrogen gas. Write the skeletal equation including these and then balance it.	$2\,Rb(s) + 2\,H_2O(l) \longrightarrow 2\,Rb^+(aq) + 2\,OH^-(aq) + H_2(g)$

—Continued on the next page

Continued—

(c) Halogens react with each other to form interhalogen compounds. Write the skeletal equation with the halogens as the reactants and the interhalogen compound as the product and balance the equation.	$Cl_2(g) + I_2(s) \longrightarrow 2\,ICl(g)$

FOR PRACTICE 4.9 Write a balanced chemical equation for each reaction.

(a) the reaction between aluminum metal and chlorine gas

(b) the reaction between lithium metal and liquid water

(c) the reaction between gaseous hydrogen and liquid bromine

QUIZ YOURSELF **NOW!**

Self-Assessment Quiz

Q1. What is the coefficient of H_2O when the reaction is balanced? **MISSED THIS?** *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

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

a) 1 b) 2
c) 3 d) 4

Q2. What are the correct coefficients (reading from left to right) when the chemical equation is balanced? **MISSED THIS?** *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

$$\_\_PCl_3(l) + \_\_H_2O(l) \longrightarrow \_\_H_3PO_3(aq) + \_\_HCl(aq)$$

a) 1, 3, 1, 3 b) 1, 2, 1, 1
c) 1, 3, 2, 1 d) 3, 6, 1, 9

Q3. For the reaction shown here, 3.5 mol A is mixed with 5.9 mol B and 2.2 mol C. What is the limiting reactant? **MISSED THIS?** *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

$$3\,A + 2\,B + C \rightarrow 2\,D$$

a) A b) B c) C d) D

Q4. Manganese(IV) oxide reacts with aluminum to form elemental manganese and aluminum oxide:

$$3\,MnO_2 + 4\,Al \longrightarrow 3\,Mn + 2\,Al_2O_3$$

What mass of Al is required to completely react with 25.0 g MnO_2? **MISSED THIS?** *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

a) 7.76 g Al b) 5.82 g Al
c) 33.3 g Al d) 10.3 g Al

Q5. Sodium and chlorine react to form sodium chloride:

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

What is the theoretical yield of sodium chloride for the reaction of 55.0 g Na with 67.2 g Cl_2? **MISSED THIS?** *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

a) 1.40×10^2 g NaCl b) 111 g NaCl
c) 55.4 g NaCl d) 222 g NaCl

Q6. Sulfur and fluorine react to form sulfur hexafluoride:

$$S(s) + 3\,F_2(g) \longrightarrow SF_6(g)$$

If 50.0 g S is allowed to react as completely as possible with 105.0 g F_2, what mass of the excess reactant is left? **MISSED THIS?** *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

a) 20.5 g S b) 45.7 g F_2
c) 15.0 g S d) 36.3 g F_2

Q7. 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? **MISSED THIS?** *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

a) 55.1% b) 44.8%
c) 123% d) 81.2%

Q8. Identify the correct balanced equation for the combustion of propane (C_3H_8). **MISSED THIS?** *Read Section 4.5*

a) $C_3H_8(g) \longrightarrow 4\,H_2(g) + 3\,C(s)$
b) $C_3H_8(g) + 5\,O_2(g) \longrightarrow 4\,H_2O(g) + 3\,CO_2(g)$
c) $C_3H_8(g) + 3\,O_2(g) \longrightarrow 4\,H_2O(g) + 3\,CO_2(g)$
d) $2\,C_3H_8(g) + 9\,O_2(g) \longrightarrow 6\,H_2CO_3(g) + 2\,H_2(g)$

Q9. Solid potassium chlorate ($KClO_3$) decomposes into potassium chloride and oxygen gas when heated. How many moles of oxygen form when 55.8 g $KClO_3$ completely decompose? **MISSED THIS?** *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

a) 0.455 mol O_2
b) 0.304 mol O_2
c) 83.7 mol O_2
d) 0.683 mol O_2

Q10. Identify the product of the reaction between hydrogen gas and bromine. **MISSED THIS?** *Read Section 4.5*

a) HBr b) H_2Br
c) HBr_2 d) H_2Br_3

Answers: **1.** (b) **2.** (a) **3.** (a) **4.** (d) **5.** (b) **6.** (a) **7.** (d) **8.** (b) **9.** (d) **10.** (a)

CHAPTER 4 IN REVIEW

TERMS

Section 4.2
chemical reaction (141)
combustion reaction (141)
chemical equation (141)
reactants (141)
products (141)
balanced chemical equation (141)

Section 4.3
stoichiometry (145)

Section 4.4
limiting reactant (150)
reactant in excess (150)
theoretical yield (150)
actual yield (150)
percent yield (150)

CONCEPTS

Climate Change and the Combustion of Fossil Fuels (4.1)

- Greenhouse gases warm Earth by trapping some of the sunlight that penetrates Earth's atmosphere. Global warming, resulting from rising atmospheric carbon dioxide levels, is potentially harmful.
- The largest atmospheric carbon dioxide source is the burning of fossil fuels. This can be verified by reaction stoichiometry.

Writing and Balancing Chemical Equations (4.2)

- In chemistry, we represent chemical reactions with chemical equations. The substances on the left-hand side of a chemical equation are the reactants, and the substances on the right-hand side are the products.
- Chemical equations are balanced when the number of each type of atom on the left side of the equation is equal to the number on the right side.

Reaction Stoichiometry (4.3)

- 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 (4.4)

- When a chemical reaction actually occurs, the reactants are usually not present in the exact stoichiometric ratios specified by the balanced chemical equation. The limiting reactant is the one that is available 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 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 when the reaction is carried out is the percent yield.

Combustion, Alkali Metals, and Halogens (4.5)

- In a combustion reaction, a substance reacts with oxygen—emitting heat and forming one or more oxygen-containing products. The alkali metals react with nonmetals, losing electrons in the process.
- The halogens react with many metals to form metal halides. They also react with hydrogen to form hydrogen halides and with one another to form interhalogen compounds.

EQUATIONS AND RELATIONSHIPS

Mass-to-Mass Conversion: Stoichiometry (4.2)

$$\text{mass A} \longrightarrow \text{amount A (in moles)} \longrightarrow \text{amount B (in moles)} \longrightarrow \text{mass B}$$

Percent Yield (4.3)

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

LEARNING OUTCOMES

Chapter Objectives	Assessment
Balance chemical equations (4.2)	Examples 4.1, 4.2, 4.3 **For Practice** 4.1, 4.2, 4.3 **Exercises** 13–24
Perform calculations involving the stoichiometry of a reaction (4.3)	Examples 4.4, 4.5 **For Practice** 4.4, 4.5 **Exercises** 25–34
Analyze chemical reactions involving a limiting reactant (4.4)	Examples 4.6, 4.7 **For Practice** 4.6, 4.7 **Exercises** 35–50
Write chemical equations for combustion reactions (4.5)	Example 4.8 **For Practice** 4.8 **Exercises** 51–52
Write chemical equations for reactions involving alkali metals and halogens (4.5)	Example 4.9 **For Practice** 4.9 **Exercises** 53–58

EXERCISES

Mastering Chemistry provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

REVIEW QUESTIONS

1. What is the greenhouse effect?

2. Why are scientists concerned about increases in atmospheric carbon dioxide? What is the source of the increase?

3. What is a balanced chemical equation?

4. Identify the reactants and products in this chemical equation.
$4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(g)$

5. Why must chemical equations be balanced?

6. What is reaction stoichiometry? What is the significance of the coefficients in a balanced chemical equation?

7. In a chemical reaction, what is the limiting reactant? What do we mean when we say a reactant is in excess?

8. In a chemical reaction, what is the theoretical yield and the percent yield?

9. We typically calculate the percent yield using the actual yield and theoretical yield in units of mass (grams or kilograms). Would the percent yield be different if the actual yield and theoretical yield were in units of amount (moles)?

10. What is a combustion reaction? Why are combustion reactions important? Give an example.

11. Write a general equation for the reaction of an alkali metal with:

a. a halogen
b. water

12. Write a general equation for the reaction of a halogen with:

a. a metal **b.** hydrogen
c. another halogen

PROBLEMS BY TOPIC

Writing and Balancing Chemical Equations

13. Sulfuric acid is a component of acid rain formed when gaseous sulfur dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous sulfuric acid. Write the balanced chemical equation for this reaction. (*Note:* This is a simplified representation of this reaction.)

MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

14. Nitric acid is a component of acid rain that forms when gaseous nitrogen dioxide pollutant reacts with gaseous oxygen and liquid water to form aqueous nitric acid. Write the balanced chemical equation for this reaction. (*Note:* This is a simplified representation of this reaction.)

15. In a popular classroom demonstration, solid sodium is added to liquid water and reacts to produce hydrogen gas and aqueous sodium hydroxide. Write the balanced chemical equation for this reaction.

MISSED THIS? *Read Sections 4.2, 4.5; Watch KCV 4.2, IWE 4.2*

16. When iron rusts, solid iron reacts with gaseous oxygen to form solid iron(III) oxide. Write the balanced chemical equation for this reaction.

17. Write the balanced chemical equation for the fermentation of sucrose ($C_{12}H_{22}O_{11}$) by yeasts in which the aqueous sugar reacts with water to form aqueous ethanol (C_2H_5OH) and carbon dioxide gas.

MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

18. Write the balanced equation for the photosynthesis reaction in which gaseous carbon dioxide and liquid water react in the presence of chlorophyll to produce aqueous glucose ($C_6H_{12}O_6$) and oxygen gas.

19. Write the balanced chemical equation for each reaction.
MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2*

a. Solid lead(II) sulfide reacts with aqueous hydrobromic acid to form solid lead(II) bromide and dihydrogen monosulfide gas.

b. Gaseous carbon monoxide reacts with hydrogen gas to form gaseous methane (CH_4) and liquid water.

c. Aqueous hydrochloric acid reacts with solid manganese(IV) oxide to form aqueous manganese(II) chloride, liquid water, and chlorine gas.

d. Liquid pentane (C_5H_{12}) reacts with gaseous oxygen to form carbon dioxide and liquid water.

20. Write the balanced chemical equation for each reaction.

a. Solid copper reacts with solid sulfur to form solid copper(I) sulfide.

b. Solid iron(III) oxide reacts with hydrogen gas to form solid iron and liquid water.

c. Sulfur dioxide gas reacts with oxygen gas to form sulfur trioxide gas.

d. Gaseous ammonia (NH_3) reacts with gaseous oxygen to form gaseous nitrogen monoxide and gaseous water.

21. Write the balanced chemical equation for the reaction of aqueous sodium carbonate with aqueous copper(II) chloride to form solid copper(II) carbonate and aqueous sodium chloride.

MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.3*

22. Write the balanced chemical equation for the reaction of aqueous potassium hydroxide with aqueous iron(III) chloride to form solid iron(III) hydroxide and aqueous potassium chloride.

23. Balance each chemical equation.

MISSED THIS? *Read Section 4.2; Watch KCV 4.2, IWE 4.2, 4.3*

a. $CO_2(g) + CaSiO_3(s) + H_2O(l) \longrightarrow SiO_2(s) + Ca(HCO_3)_2(aq)$
b. $Co(NO_3)_3(aq) + (NH_4)_2S(aq) \longrightarrow Co_2S_3(s) + NH_4NO_3(aq)$
c. $Cu_2O(s) + C(s) \longrightarrow Cu(s) + CO(g)$
d. $H_2(g) + Cl_2(g) \longrightarrow HCl(g)$

24. Balance each chemical equation.

a. $Na_2S(aq) + Cu(NO_3)_2(aq) \longrightarrow NaNO_3(aq) + CuS(s)$
b. $N_2H_4(l) \longrightarrow NH_3(g) + N_2(g)$
c. $HCl(aq) + O_2(g) \longrightarrow H_2O(l) + Cl_2(g)$
d. $FeS(s) + HCl(aq) \longrightarrow FeCl_2(aq) + H_2S(g)$

Reaction Stoichiometry

25. 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 7.2 moles of C_6H_{14}.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

26. 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 0.461 mole of $HC_2H_3O_2$.

27. Calculate how many moles of NO_2 form when each quantity of reactant completely reacts.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

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

a. 2.5 mol N_2O_5
b. 6.8 mol N_2O_5
c. 15.2 g N_2O_5
d. 2.87 kg N_2O_5

28. Calculate how many moles of NH_3 form when each quantity of reactant completely reacts.

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

a. 2.6 mol N_2H_4 **b.** 3.55 mol N_2H_4
c. 65.3 g N_2H_4 **d.** 4.88 kg N_2H_4

29. Consider the balanced equation:

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

Complete the table showing the appropriate number of moles of reactants and products. If 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 that forms. If 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 forms.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

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

30. Consider the balanced equation:

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

Complete the table showing the appropriate number of moles of reactants and products. If 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 that forms. If 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 forms.

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

31. Hydrobromic acid dissolves solid iron according to the reaction:

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

What mass of HBr (in g) do you need to dissolve a 3.2-g pure iron bar on a padlock? What mass of H_2 would the complete reaction of the iron bar produce?

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

32. Sulfuric acid dissolves aluminum metal according to the reaction:

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

Suppose you want to dissolve an aluminum block with a mass of 15.2 g. What minimum mass of H_2SO_4 (in g) do you need? What mass of H_2 gas (in g) does the complete reaction of the aluminum block produce?

33. For each of the reactions, calculate the mass (in grams) of the product that forms when 3.67 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

MISSED THIS? *Read Section 4.3; Watch KCV 4.3, IWE 4.4*

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

34. For each of the reactions, calculate the mass (in grams) of the product that forms when 15.39 g of the underlined reactant completely reacts. Assume that there is more than enough of the other reactant.

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

Limiting Reactant, Theoretical Yield, and Percent Yield

35. Find the limiting reactant for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

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

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

36. Find the limiting reactant for each initial amount of reactants.

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

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

37. Consider the reaction:

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

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

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

a. b. c.

38. Consider the reaction:

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

Each of the molecular diagrams represents an initial mixture of the reactants. Which reaction mixture produces the greatest amount of products? How many CO_2 molecules form from the reaction mixture that produces the greatest amount of products?

a. b. c.

39. Calculate the theoretical yield of the product (in moles) for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

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

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

40. Calculate the theoretical yield of product (in moles) for each initial amount of reactants.

$$3\,Mn(s) + 2\,O_2(g) \longrightarrow Mn_3O_4(s)$$

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

41. Zinc sulfide reacts with oxygen according to the reaction:

$$2\,ZnS(s) + 3\,O_2(g) \longrightarrow 2\,ZnO(s) + 2\,SO_2(g)$$

A reaction mixture initially contains 4.2 mol ZnS and 6.8 mol O_2. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant remains?

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

42. Iron(II) sulfide reacts with hydrochloric acid according to the reaction:

$$FeS(s) + 2\,HCl(aq) \longrightarrow FeCl_2(s) + H_2S(g)$$

A reaction mixture initially contains 0.223 mol FeS and 0.652 mol HCl. Once the reaction has occurred as completely as possible, what amount (in moles) of the excess reactant remains?

43. For the reaction shown, calculate the theoretical yield of product (in grams) for each initial amount of reactants.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

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

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

44. For the reaction shown, calculate the theoretical yield of the product (in grams) for each initial amount of reactants.

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

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

45. Iron(III) oxide reacts with carbon monoxide according to the equation:

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

A reaction mixture initially contains 22.55 g Fe_2O_3 and 14.78 g CO. Once the reaction has occurred as completely as possible, what mass (in g) of the excess reactant remains?

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

46. Elemental phosphorus reacts with chlorine gas according to the equation:

$$P_4(s) + 6\,Cl_2(g) \longrightarrow 4\,PCl_3(l)$$

A reaction mixture initially contains 45.69 g P_4 and 131.3 g Cl_2. Once the reaction has occurred as completely as possible, what mass (in g) of the excess reactant remains?

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

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

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 and 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.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

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

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

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.

49. Urea (CH_4N_2O) is a common fertilizer that is synthesized by the reaction of ammonia (NH_3) with carbon dioxide:

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

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.

MISSED THIS? *Read Section 4.4; Watch KCV 4.4, IWE 4.6*

50. 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.

Combustion, Alkali Metal, and Halogen Reactions

51. Complete and balance each combustion reaction equation.

MISSED THIS? *Read Section 4.5*

a. $S(s) + O_2(g) \longrightarrow$ **b.** $C_3H_6(g) + O_2(g) \longrightarrow$
c. $Ca(s) + O_2(g) \longrightarrow$ **d.** $C_5H_{12}S(l) + O_2(g) \longrightarrow$

52. Complete and balance each combustion reaction equation:

a. $C_4H_6(g) + O_2(g) \longrightarrow$ **b.** $C(s) + O_2(g) \longrightarrow$
c. $CS_2(s) + O_2(g) \longrightarrow$ **d.** $C_3H_8O(l) + O_2(g) \longrightarrow$

53. Write a balanced chemical equation for the reaction of solid strontium with iodine gas.

MISSED THIS? *Read Section 4.5*

54. Write a balanced chemical equation for the reaction between lithium metal and chlorine gas.

55. Write a balanced chemical equation for the reaction of solid lithium with liquid water. **MISSED THIS?** *Read Section 4.5*

56. Write a balanced chemical equation for the reaction of solid potassium with liquid water.

57. Write a balanced equation for the reaction of hydrogen gas with bromine gas. **MISSED THIS?** *Read Section 4.5*

58. Write a balanced equation for the reaction of chlorine gas with fluorine gas.

CUMULATIVE PROBLEMS

59. 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.)

60. 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.)

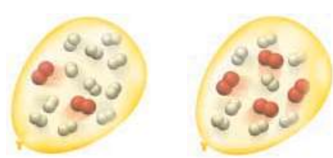
61. Aspirin can be made in the laboratory by reacting acetic anhydride ($C_4H_6O_3$) with salicylic acid ($C_7H_6O_3$) to form aspirin ($C_9H_8O_4$) and acetic acid ($C_2H_4O_2$). The balanced equation is:

$$C_4H_6O_3 + C_7H_6O_3 \longrightarrow C_9H_8O_4 + C_2H_4O_2$$

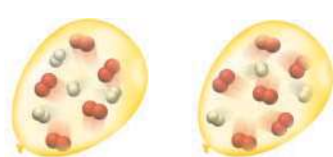
In a laboratory synthesis, a student begins with 3.00 mL of acetic anhydride (density = 1.08 g/mL) and 1.25 g of salicylic acid. Once the reaction is complete, the student collects 1.22 g of aspirin. Determine the limiting reactant, theoretical yield of aspirin, and percent yield for the reaction.

62. 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.)

63. 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 only with hydrogen—the intensity of the explosion depends on the relative amounts of oxygen and hydrogen within the balloon. Look at the molecular views representing different amounts of hydrogen and oxygen in four different balloons. Based on the balanced chemical equation, which balloon will make the loudest explosion?

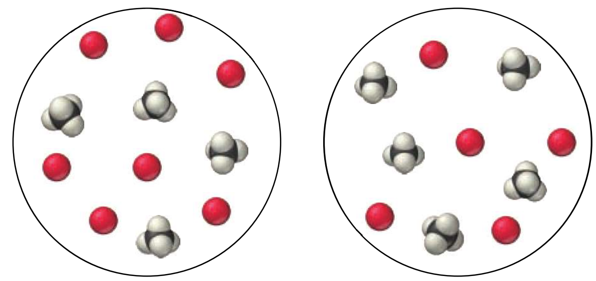


a. b.

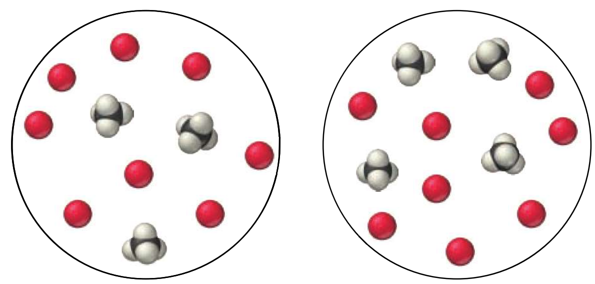


c. d.

64. Gaseous methane reacts with oxygen to form carbon dioxide and water vapor. Write a balanced equation for the combustion reaction and determine which mixture has neither reactant in excess.



a. b.



c. d.

65. The reaction of NH_3 and O_2 forms NO and water. The NO can be used to convert P_4 to P_4O_6, forming N_2 in the process. The P_4O_6 can be treated with water to form H_3PO_3, which forms PH_3 and H_3PO_4 when heated. Find the mass of PH_3 that forms from the reaction of 1.00 g of NH_3.

66. An important reaction that takes place in a blast furnace during the production of iron is the formation of iron metal and CO_2 from Fe_2O_3 and CO. Determine the mass of Fe_2O_3 required to form 910 kg of iron. Determine the amount of CO_2 that forms in this process.

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

68. Titanium occurs in the magnetic mineral ilmenite ($FeTiO_3$), which is often found mixed with sand. The ilmenite can be separated from the sand with magnets. The titanium can then be extracted from the ilmenite by the following set of reactions:

$$FeTiO_3(s) + 3\,Cl_2(g) + 3\,C(s) \longrightarrow 3\,CO(g) + FeCl_2(s) + TiCl_4(g)$$

$$TiCl_4(g) + 2\,Mg(s) \longrightarrow 2\,MgCl_2(l) + Ti(s)$$

Suppose that an ilmenite–sand mixture contains 22.8% ilmenite by mass and that 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

69. A mixture of C_3H_8 and C_2H_2 has a mass of 2.0 g. It is burned in excess O_2 to form a mixture of water and carbon dioxide that contains 1.5 times as many moles of CO_2 as of water. Find the mass of C_2H_2 in the original mixture.

70. A mixture of 20.6 g of P and 79.4 g of Cl_2 reacts completely to form PCl_3 and PCl_5 as the only products. Find the mass of PCl_3 that forms.

71. A mixture of A and B contains a total of 5.3 mols. Both A and B react with Z according to the following equations:

$$A + Z \longrightarrow AZ$$
$$B + 2\,Z \longrightarrow BZ_2$$

The reaction of the mixture of A and B with Z consumes 7.8 mol Z. Assuming the reactions go to completion, how many moles of A does the mixture contain?

72. 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 following chemical reaction:

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

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 given 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.)

73. 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.

74. 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

75. Consider the reaction:

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

The molar mass of K is 39.10 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

76. Consider the reaction:

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

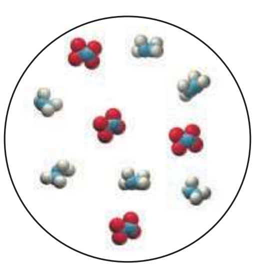
A reaction mixture initially contains 5 moles of NO and 10 moles of H_2. Without doing any calculations, determine which set of amounts best represents the mixture after the reactants have reacted as completely as possible. Explain your reasoning.

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

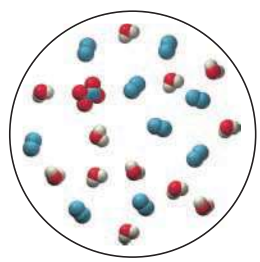
77. Consider the reaction:

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

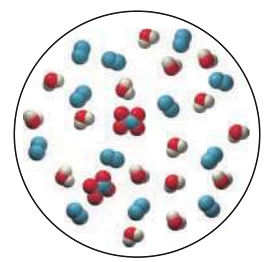
Consider also this representation of an initial mixture of N_2H_4 and N_2O_4:



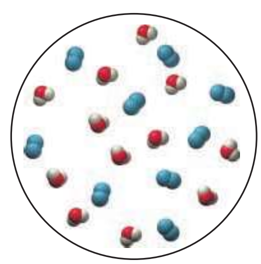
Which diagram best represents the reaction mixture after the reactants have reacted as completely as possible?



a.



b.



c.

QUESTIONS FOR GROUP WORK

Active Classroom Learning

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

78. Octane (C_8H_{18}), a component of gasoline, reacts with oxygen to form carbon dioxide and water. Write the balanced chemical reaction for this process by passing a single piece of paper around your group and asking each group member to complete the next logical step. As each member completes his or her step, explain your reasoning to the group.

79. Imagine you mix 16.05 g of methane (CH_4) gas and 96.00 g of oxygen (O_2) gas and then ignite the mixture. After a bright flash and a loud bang, some water vapor forms.

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

b. Depict the process that occurred 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 circles bonded together) represent exactly one mole.

c. How many moles of water can you make? How many moles of carbon dioxide?

d. Will anything be left over? If so, how much?

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

80. Explain the problem with the following statement to your group and correct it. "When a chemical equation is balanced, the number of molecules of each type on both sides of the equation is equal."

DATA INTERPRETATION AND ANALYSIS

Limiting Reactant and Percent Yield

81. A chemical reaction in which reactants A and B form the product C is studied in the laboratory. The researcher carries out the reaction with differing relative amounts of reactants and measures the amount of product produced. Examine the given tabulated data from the experiment and answer the questions.

Experiment #	Mass A (g)	Mass B (g)	Mass C Obtained (g)
1	2.51	7.54	3.76
2	5.03	7.51	7.43
3	7.55	7.52	11.13
4	12.53	7.49	14.84
5	15.04	7.47	14.94
6	19.98	7.51	15.17
7	20.04	9.95	19.31
8	20.02	12.55	24.69

a. For which experiments is A the limiting reactant?

b. For which experiments is B the limiting reactant?

c. The molar mass of A is 50.0 g/mol, and the molar mass of B is 75.0 g/mol. What are the coeffecients of A and B in the balanced chemical equation?

d. For each of the experiments in which A is the limiting reactant, calculate the mass of B remaining after the reaction has gone to completion. Use the molar masses and coeffecients from part c.

e. The molar mass of C is 88.0 g/mol. What is the coefficient of C in the balanced chemical equation?

f. Calculate an average percent yield for the reaction.

Cc ANSWERS TO CONCEPTUAL CONNECTIONS

Counting Atoms in a Chemical Equation

4.1 **(d)** The number of oxygen atoms in 2 Fe_2O_3 is 6, and the number in 4 CO_2 is 8, for a total of 14.

Balanced Chemical Equations

4.2 **(a)** When the equation is balanced, the number of atoms of each type is the same on both sides of the equation. Since molecules change during a chemical reaction, the number of molecules is not the same on both sides, nor is the number of moles of each molecule necessarily the same.

Stoichiometry I

4.3 **(d)** $22.0\ \cancel{\text{mol } C_8H_{18}} \times \dfrac{18\text{ mol } H_2O}{2\ \cancel{\text{mol } C_8H_{18}}} = 198\text{ mol } H_2O$

Stoichiometry II

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

Stoichiometry III

4.5 **(a)** Since the balanced equation indicates that one A reacts with three B, and since the diagram in the question includes 12 B available to react, the amount of A necessary is four A.

Limiting Reactant and Theoretical Yield

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

Reactant in Excess

4.7 **(c)** 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.