Chapter 6

Chemical reactions and quantities

Figure 6.1 Many modern rocket fuels are solid mixtures of substances combined in carefully measured amounts and ignited to yield a thrust-generating chemical reaction. (credit: modification of work by NASA)

Chapter Outline

6.1 Writing and Balancing Chemical Equations

6.2 Classifying Chemical Reactions

6.3 Formula Mass and the Mole Concept

6.4 Determining Empirical and Molecular Formulas

6.5 Reaction Stoichiometry

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Introduction

Solid-fuel rockets are a central feature in the world's space exploration programs, including the new Space Launch System being developed by the National Aeronautics and Space Administration (NASA) to replace the retired Space Shuttle fleet (**Figure 6.1**). The engines of these rockets rely on carefully prepared solid mixtures of chemicals combined in precisely measured amounts. Igniting the mixture initiates a vigorous chemical reaction that rapidly generates large amounts of gaseous products. These gases are ejected from the rocket engine through its nozzle, providing the thrust needed to propel heavy payloads into space. Both the nature of this chemical reaction and the relationships between the amounts of the substances being consumed and produced by the reaction are critically important considerations that determine the success of the technology. This chapter will describe how to symbolize

chemical reactions using chemical equations, how to classify some common chemical reactions by identifying patterns of reactivity, and how to determine the quantitative relations between the amounts of substances involved in chemical reactions—that is, the reaction *stoichiometry*.

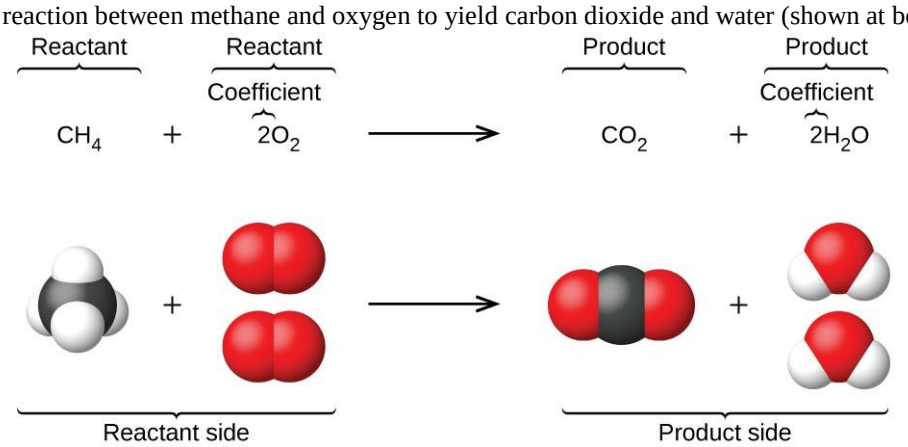
6.1 Writing and Balancing Chemical Equations

By the end of this section, you will be able to:

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations in molecular, total ionic, and net ionic formats.

The preceding chapter introduced the use of element symbols to represent individual atoms. When atoms gain or lose electrons to yield ions, or combine with other atoms to form molecules, their symbols are modified or combined to generate chemical formulas that appropriately represent these species. Extending this symbolism to represent both the identities and the relative quantities of substances undergoing a chemical (or physical) change involves writing and balancing a **chemical equation**. Consider as an example the reaction between one methane molecule (CH_4) and two diatomic oxygen molecules (O_2) to produce one carbon dioxide molecule (CO_2) and two water molecules (H_2O). The chemical equation representing this process is provided in the upper half of **Figure 6.2**, with space-filling molecular models shown in the lower half of the figure.

Figure 6.2 The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be



represented by a chemical equation using formulas (top).

This example illustrates the fundamental aspects of any chemical equation:

1. The substances undergoing reaction are called **reactants**, and their formulas are placed on the left side of the equation.
2. The substances generated by the reaction are called **products**, and their formulas are placed on the right side of the equation.
3. Plus signs (+) separate individual reactant and product formulas, and an arrow (⟶) separates the reactant and product (left and right) sides of the equation.

4. The relative numbers of reactant and product species are represented by **coefficients** (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

It is common practice to use the smallest possible whole-number coefficients in a chemical equation, as is done in this example. Realize, however, that these coefficients represent the *relative* numbers of reactants and products, and, therefore, they may be correctly interpreted as ratios. Methane and oxygen react to yield carbon dioxide and water in a 1:2:1:2 ratio. This ratio is satisfied if the numbers of these molecules are, respectively, 1-2-1-2, or 2-4-2-4, or 3-6-3-6, and so on (**Figure 6.3**). Likewise, these coefficients may be interpreted with regard to any amount (number) unit, and so this equation may be correctly read in many ways, including:

- *One* methane molecule and *two* oxygen molecules react to yield *one* carbon dioxide molecule and *two* water molecules.
- *One dozen* methane molecules and *two dozen* oxygen molecules react to yield *one dozen* carbon dioxide molecules and *two dozen* water molecules.
- *One mole* of methane molecules and *2 moles* of oxygen molecules react to yield *1 mole* of carbon dioxide molecules and *2 moles* of water molecules.

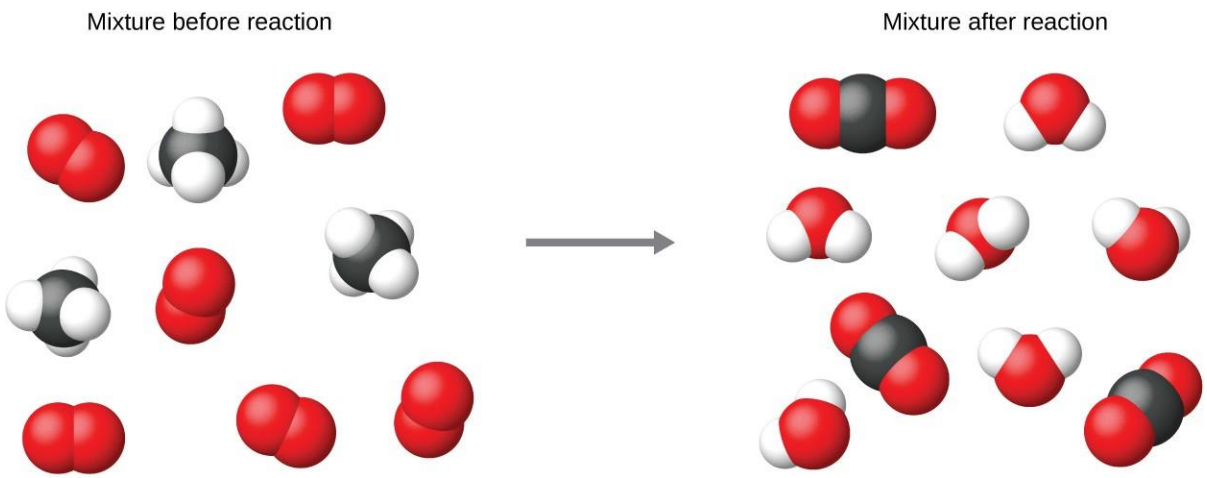


Figure 6.3 Regardless of the absolute numbers of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given by the chemical reaction equation.

Balancing Equations

The chemical equation described in section 4.1 is **balanced**, meaning that equal numbers of atoms for each element involved in the reaction are represented on the reactant and product sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example

reaction, CO_2 and H_2O, contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

$$\left(1\ CO_2 molecule * \frac{20\ atoms}{CO_2\ molecule}\right) + \left(2H_2O\ molecules * \frac{10\ atom}{H_2O\ molecule}\right) = 40\ atoms$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:

$$\mathrm{CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O}$$

Element	Reactants	Products	Balanced?
C	1 × 1 = 1	1 × 1 = 1	1 = 1, yes
H	4 × 1 = 4	2 × 2 = 4	4 = 4, yes
O	2 × 2 = 4	(1 × 2) + (2 × 1) = 4	4 = 4, yes

A balanced chemical equation often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an *unbalanced* chemical equation:

$$\mathrm{H_2O \longrightarrow H_2 + O_2} \qquad \text{(unbalanced)}$$

Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

Element	Reactants	Products	Balanced?
H	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 1 = 1	1 × 2 = 2	1 ≠ 2, no

The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the *coefficients* of the equation may be changed as needed. Keep in mind, of course, that the *formula subscripts* define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For example, changing the reactant formula from H_2O to H_2O_2 would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for H_2O to 2.

$$\mathbf{2}\mathrm{H_2O \longrightarrow H_2 + O_2} \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
H	**2** × 2 = 4	1 × 2 = 2	4 ≠ 2, no
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the H_2 product to 2.

$2H_2O \longrightarrow \mathbf{2}H_2 + O_2$ (balanced)

Element	Reactants	Products	Balanced?
H	2 × 2 = 4	**2** × 2 = 4	4 = 4, yes
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides, and the balanced equation is, therefore:

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

Example 6.1

Balancing Chemical Equations: Write a balanced equation for the reaction of molecular nitrogen (N_2) and oxygen (O_2) to form dinitrogen pentoxide.

Solution:

First, write the unbalanced equation.

$N_2 + O_2 \longrightarrow N_2O_5$ (unbalanced)

Next, count the number of each type of atom present in the unbalanced equation.

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 2 = 2	1 × 5 = 5	2 ≠ 5, no

Though nitrogen is balanced, changes in coefficients are needed to balance the number of oxygen atoms. To balance the number of oxygen atoms, a reasonable first attempt would be to change the coefficients for the O_2 and N_2O_5 to integers that will yield 10 O atoms (the least common multiple for the O atom subscripts in these two formulas).

$N_2 + \mathbf{5}O_2 \longrightarrow \mathbf{2}N_2O_5$ (unbalanced)

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	**2** × 2 = 4	2 ≠ 4, no
O	**5** × 2 = 10	**2** × 5 = 10	10 = 10, yes

The N atom balance has been upset by this change; it is restored by changing the coefficient for the reactant N_2 to 2.

$$2N_2 + 5O_2 \longrightarrow 2N_2O_5$$

Element	Reactants	Products	Balanced?
N	**2** × 2 = 4	2 × 2 = 4	4 = 4, yes
O	5 × 2 = 10	2 × 5 = 10	10 = 10, yes

The numbers of N and O atoms on either side of the equation are now equal, and so the equation is balanced.

Check Your Learning

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen, and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)

Answer: $2NH_4NO_3 \longrightarrow 2N_2 + O_2 + 4H_2O$

It is sometimes convenient to use fractions instead of integers as intermediate coefficients in the process of balancing a chemical equation. When balance is achieved, all the equation's coefficients may then be multiplied by a whole number to convert the fractional coefficients to integers without upsetting the atom balance. For example, consider the reaction of ethane (C_2H_6) with oxygen to yield H_2O and CO_2, represented by the unbalanced equation:

$$C_2H_6 + O_2 \longrightarrow H_2O + CO_2 \qquad \text{(unbalanced)}$$

Following the usual inspection approach, one might first balance C and H atoms by changing the coefficients for the two product species, as shown:

$$C_2H_6 + O_2 \longrightarrow 3H_2O + 2CO_2 \qquad \text{(unbalanced)}$$

This results in seven O atoms on the product side of the equation, an odd number—no integer coefficient can be used

with the O_2 reactant to yield an odd number, so a fractional coefficient, $\frac{7}{2}$, is used instead to yield a provisional balanced equation:

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 3H_2O + 2CO_2$$

A conventional balanced equation with integer-only coefficients is derived by multiplying each coefficient by 2:

$$2C_2H_6 + 7O_2 \longrightarrow 6H_2O + 4CO_2$$

Finally, with regard to balanced equations, recall that convention dictates use of the *smallest whole-number coefficients*. Although the equation for the reaction between molecular nitrogen and molecular hydrogen to produce ammonia is, indeed, balanced,

$$3N_2 + 9H_2 \longrightarrow 6NH_3$$

the coefficients are not the smallest possible integers representing the relative numbers of reactant and product molecules. Dividing each coefficient by the greatest common factor, 3, gives the preferred equation:

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

Link to Learning

Use this interactive **tutorial (http://openstaxcollege.org/l/16BalanceEq)** for additional practice balancing equations.

Additional Information in Chemical Equations

The physical states of reactants and products in chemical equations very often are indicated with a parenthetical abbreviation following the formulas. Common abbreviations include for solids, *l* for liquids, *g* for gases, and *aq for* substances dissolved in water (*aqueous solutions*, as introduced in the preceding chapter). These notations are illustrated in the example equation here:

$$2\mathrm{Na}(s) + 2\mathrm{H_2O}(l) \longrightarrow 2\mathrm{NaOH}(aq) + \mathrm{H_2}(g)$$

This equation represents the reaction that takes place when sodium metal is placed in water. The solid sodium reacts with liquid water to produce molecular hydrogen gas and the ionic compound sodium hydroxide (a solid in pure form, but readily dissolved in water).

Special conditions necessary for a reaction are sometimes designated by writing a word or symbol above or below the equation's arrow. For example, a reaction carried out by heating may be indicated by the uppercase Greek letter delta (Δ) over the arrow.

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

Other examples of these special conditions will be encountered in more depth in later chapters.

6.2 Classifying Chemical Reactions

By the end of this section, you will be able to:

- Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction)
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Identify common acids and bases
- Predict the solubility of common inorganic compounds by using solubility rules
- Compute the oxidation states for elements in compounds

Humans interact with one another in various and complex ways, and we classify these interactions according to common patterns of behavior. When two humans exchange information, we say they are communicating. When they exchange blows with their fists or feet, we say they are fighting. Faced with a wide range of varied interactions between

chemical substances, scientists have likewise found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This module will provide an introduction to three of the most prevalent types of chemical reactions: precipitation, acid-base, and oxidation-reduction.

Oxidation-Reduction Reactions

Earth's atmosphere contains about 20% molecular oxygen, O_2, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape the world. The term **oxidation** was originally used to describe chemical reactions involving O_2, but its meaning has evolved to refer to a broad and important reaction class known as *oxidation-reduction (redox) reactions*. A few examples of such reactions will be used to develop a clear picture of this classification.

Some redox reactions involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride:

$$2\mathrm{Na}(s) + \mathrm{Cl}_2(g) \longrightarrow 2\mathrm{NaCl}(s)$$

It is helpful to view the process with regard to each individual reactant, that is, to represent the fate of each reactant in the form of an equation called a **half-reaction**:

$$2\mathrm{Na}(s) \longrightarrow 2\mathrm{Na}^+(s) + 2\mathrm{e}^- \quad \mathrm{Cl}_2(g) + 2\mathrm{e}^- \longrightarrow 2\mathrm{Cl}^-(s)$$

These equations show that Na atoms *lose electrons* while Cl atoms (in the Cl_2 molecule) *gain electrons*, the "*s*" subscripts for the resulting ions signifying they are present in the form of a solid ionic compound. For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur:

oxidation = loss of electrons

reduction = gain of electrons

In this reaction, then, sodium is *oxidized*, and chlorine undergoes **reduction**. Viewed from a more active perspective, sodium functions as a **reducing agent (reductant)**, since it provides electrons to (or reduces) chlorine. Likewise, chlorine functions as an **oxidizing agent (oxidant)**, as it effectively removes electrons from (oxidizes) sodium.

reducing agent = species that is oxidized

oxidizing agent = species that is reduced

Some redox processes, however, do not involve the transfer of electrons. Consider, for example, a reaction similar to the one yielding NaCl:

$$\mathrm{H}_2(g) + \mathrm{Cl}_2(g) \longrightarrow 2\mathrm{HCl}(g)$$

The product of this reaction is a covalent compound, so transfer of electrons in the explicit sense is not involved. To clarify the similarity of this reaction to the previous one and permit an unambiguous definition of redox reactions, a property called *oxidation number* has been defined. The **oxidation number** (or **oxidation state**) of an element in a compound is the charge its atoms would possess *if the compound was ionic*. The following guidelines are used to assign oxidation numbers to each element in a molecule or ion.

1. The oxidation number of an atom in an elemental substance is zero.

2. The oxidation number of a monatomic ion is equal to the ion's charge.

3. Oxidation numbers for common nonmetals are usually assigned as follows:

- Hydrogen: +1 when combined with nonmetals, −1 when combined with metals
- Oxygen: −2 in most compounds, sometimes −1 (so-called peroxides, O_2^{2-}), very rarely $-\frac{1}{2}$ (so-called superoxides, O_2^{2-}), positive values when combined with F (values vary) Halogens: −1 for F always, −1 for other halogens except when combined with oxygen or other halogens (positive oxidation numbers in these cases, varying values)

4. The sum of oxidation numbers for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion.

Note: The proper convention for reporting charge is to write the number first, followed by the sign (e.g., 2+), while oxidation number is written with the reversed sequence, sign followed by number (e.g., +2). This convention aims to emphasize the distinction between these two related properties.

Example 6.2

Assigning Oxidation Numbers: Follow the guidelines in this section of the text to assign oxidation numbers to all the elements in the following species:

(a) H_2S

(b) SO_3^{2-}

(c) Na_2SO_4

Solution:

(a) According to guideline 1, the oxidation number for H is +1.

Using this oxidation number and the compound's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$charge\ on\ H_2S = 0 = (2 * +1) + (1 * x)$$
$$X = 0 - (2 * +1) = -2$$

(b) Guideline 3 suggests the oxidation number for oxygen is −2.

Using this oxidation number and the ion's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$Charge\ on\ SO_3^{-2} = -2 = (3 * -2) + (1 * X)$$
$$X = -2 - (3 * -2) = +4$$

(c) For ionic compounds, it's convenient to assign oxidation numbers for the cation and anion separately. According to guideline 2, the oxidation number for sodium is +1.

Assuming the usual oxidation number for oxygen (−2 per guideline 3), the oxidation number for sulfur is calculated

as directed by guideline 4:

$$charge\ on\ SO_4^{-2} = -2 = (4 * -2) + (1 * x)$$
$$x = -2 - (4 * -2) = +6$$

Check Your Learning

Assign oxidation states to the elements whose atoms are underlined in each of the following compounds or ions:

(a) KNO3

(b) AlH3

(c) NH_4^+

(d) $H_2PO_4^-$

Answer: (a) N, +5; (b) Al, +3; (c) N, −3; (d) P, +5

Using the oxidation number concept, an all-inclusive definition of redox reaction has been established. **Oxidation-reduction (redox) reactions** are those in which one or more elements involved undergo a change in oxidation number. (While the vast majority of redox reactions involve changes in oxidation number for two or more elements, a few interesting exceptions to this rule do exist **Example 6.3**.) Definitions for the complementary processes of this reaction class are correspondingly revised as shown here:

oxidation = increase in oxidation number

reduction = decrease in oxidation number

Returning to the reactions used to introduce this topic, they may now both be identified as redox processes. In the reaction between sodium and chlorine to yield sodium chloride, sodium is oxidized (its oxidation number increases from 0 in Na to +1 in NaCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in NaCl). In the reaction between molecular hydrogen and chlorine, hydrogen is oxidized (its oxidation number increases from 0 in H_2 to +1 in HCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in HCl).

Several subclasses of redox reactions are recognized, including **combustion reactions** in which the reductant (also called a *fuel*) and oxidant (often, but not necessarily, molecular oxygen) react vigorously and produce significant amounts of heat, and often light, in the form of a flame. Solid rocket-fuel reactions such as the one depicted in **Figure 6.1** are combustion processes. A typical propellant reaction in which solid aluminum is oxidized by ammonium perchlorate is represented by this equation:

$$10Al(s) + 6NH_4ClO_4(s) \longrightarrow 4Al_2O_3(s) + 2AlCl_3(s) + 12H_2O(g) + 3N_2(g)$$

Link to Learning

Watch a brief **video (http://openstaxcollege.org/l/16hybridrocket)** showing the test firing of a small-scale, prototype, hybrid rocket engine planned for use in the new Space Launch System being developed by NASA. The first engines firing at 3 s (green flame) use a liquid fuel/oxidant mixture, and the second, more powerful engines firing at 4 s (yellow flame) use a solid mixture.

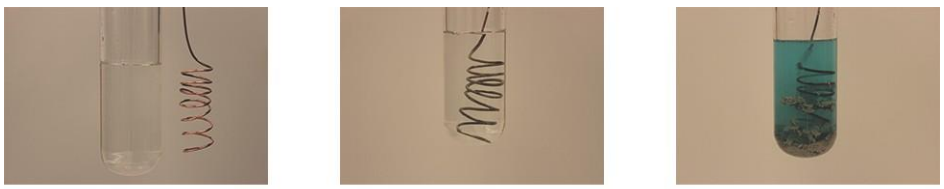
Single-displacement (replacement) reactions are redox reactions in which an ion in solution is displaced (or replaced) via the oxidation of a metallic element. One common example of this type of reaction is the acid oxidation of certain metals:

$$\mathrm{Zn}(s) + 2\mathrm{HCl}(aq) \longrightarrow \mathrm{ZnCl_2}(aq) + \mathrm{H_2}(g)$$

Metallic elements may also be oxidized by solutions of other metal salts; for example:

$$\mathrm{Cu}(s) + 2\mathrm{AgNO_3}(aq) \longrightarrow \mathrm{Cu\ (NO_3)_2}(aq) + 2\mathrm{Ag}(s)$$

This reaction may be observed by placing copper wire in a solution containing a dissolved silver salt. Silver ions in solution are reduced to elemental silver at the surface of the copper wire, and the resulting Cu^{2+} ions dissolve in the solution to yield a characteristic blue color (**Figure 6.8**).



(a) (b) (c)

Figure 6.8 (a) A copper wire is shown next to a solution containing silver(I) ions. (b) Displacement of dissolved silver ions by copper ions results in (c) accumulation of gray-colored silver metal on the wire and development of a blue color in the solution, due to dissolved copper ions. (credit: modification of work by Mark Ott)

Example 6.3

Describing Redox Reactions: Identify which equations represent redox reactions, providing a name for the reaction if appropriate. For those reactions identified as redox, name the oxidant and reductant.

(a) $\mathrm{ZnCO_3}(s) \longrightarrow \mathrm{ZnO}(s) + \mathrm{CO_2}(g)$

(b) $2\mathrm{Ga}(l) + 3\mathrm{Br_2}(l) \longrightarrow 2\mathrm{GaBr_3}(s)$

(c) $2\mathrm{H_2O_2}(aq) \longrightarrow 2\mathrm{H_2O}(l) + \mathrm{O_2}(g)$

(d) $\mathrm{BaCl_2}(aq) + \mathrm{K_2SO_4}(aq) \longrightarrow \mathrm{BaSO_4}(s) + 2\mathrm{KCl}(aq)$

(e) $\mathrm{C_2H_4}(g) + 3\mathrm{O_2}(g) \longrightarrow 2\mathrm{CO_2}(g) + 2\mathrm{H_2O}(l)$

Solution:

Redox reactions are identified per definition if one or more elements undergo a change in oxidation number.

(a) This is not a redox reaction, since oxidation numbers remain unchanged for all elements.

(b) This is a redox reaction. Gallium is oxidized, its oxidation number increasing from 0 in Ga(*l*) to +3 in $GaBr_3$(*s*). The reducing agent is Ga(*l*). Bromine is reduced, its oxidation number decreasing from 0 in Br_2(*l*) to −1 in $GaBr_3$(*s*). The oxidizing agent is Br_2(*l*).

(c) This is a redox reaction. It is a particularly interesting process, as it involves the same element, oxygen, undergoing both oxidation and reduction (a so-called *disproportionation reaction)*. Oxygen is oxidized, its oxidation number increasing from −1 in H_2O_2(*aq*) to 0 in O_2(*g*). Oxygen is also reduced, its oxidation number decreasing from

−1 in $H_2O_2(aq)$ to −2 in $H_2O(l)$. For disproportionation reactions, the same substance functions as an oxidant and a reductant.

(d) This is not a redox reaction, since oxidation numbers remain unchanged for all elements. A final check of atom and charge balance confirms the equation is balanced.

(e) This is a redox reaction (combustion). Carbon is oxidized, its oxidation number increasing from −2 in C2H4(g) to +4 in CO2(g). The reducing agent (fuel) is C2H4(g). Oxygen is reduced, its oxidation number decreasing from 0 in O2(g) to −2 in H2O(l). The oxidizing agent is O2(g).

Check Your Learning

This equation describes the production of tin(II) chloride:

$$Sn(s) + 2HCl(g) \longrightarrow SnCl_2(s) + H_2(g)$$

Is this a redox reaction? If so, provide a more specific name for the reaction if appropriate, and identify the oxidant and reductant.

Answer: Yes, a single-replacement reaction. Sn(s)is the reductant, HCl(g) is the oxidant.

Check Your Learning

In acidic solution, hydrogen peroxide reacts with Fe2+ to produce Fe3+ and H2O. Write a balanced equation for this reaction.

Answer: $H_2O_2(aq) + 2H+(aq) + 2Fe^{2+} \longrightarrow 2H_2O(l) + 2Fe^{3+}$

6.3 Formula Mass and the Mole Concept

By the end of this section, you will be able to:

- Calculate formula masses for covalent and ionic compounds
- Define the amount unit mole and the related quantity Avogadro's number Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another

We can argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Dalton's atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances.

Formula Mass

In an earlier chapter, we described the development of the atomic mass unit, the concept of average atomic masses, and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the **formula mass** of a substance by summing the average atomic masses of all the atoms represented in the substance's formula.

Formula Mass for Covalent Substances

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. **Figure 6.9** outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37

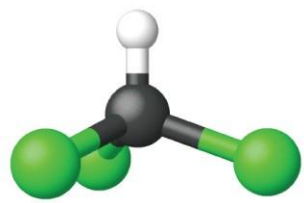


Figure 6.9 The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

Likewise, the molecular mass of an aspirin molecule, $C_9H_8O_4$, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms, and four oxygen atoms, which amounts to 180.15 amu (**Figure 6.10**).

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	9	×	12.01	=	108.09
H	8	×	1.008	=	8.064
O	4	×	16.00	=	64.00
				Molecular mass	180.15

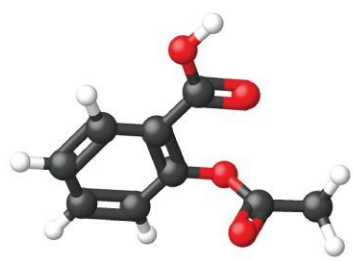


Figure 6.10 The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $C_9H_8O_4$.

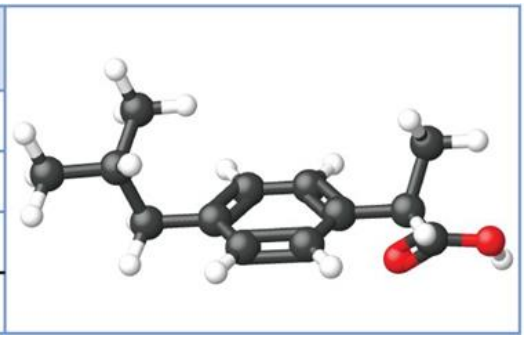
Example 6.4

Computing Molecular Mass for a Covalent Compound: Ibuprofen, $C_{13}H_{18}O_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound?

Solution:

Molecules of this compound are comprised of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.144
O	2	×	16.00	=	32.00
				Molecular mass	206.27



Check Your Learning

Acetaminophen, $C_8H_9NO_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound?

Answer: 151.16 amu

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (see **Figure 6.11**).

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

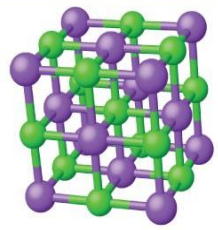


Figure 6.11 Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a

chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses.

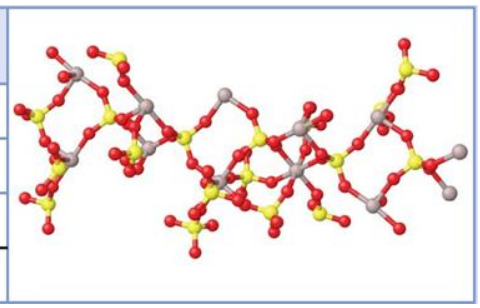
Example 6.5

Computing Formula Mass for an Ionic Compound: Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution:

The formula for this compound indicates it contains Al^{3+} and SO^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Molecular mass	342.14



Check Your Learning

Calcium phosphate, $Ca_3(PO_4)_2$, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate?

Answer: 310.18 amu

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O, and hydrogen peroxide, H_2O_2, are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the *mole*, which remains indispensable in modern chemical science.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. A **mole** is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure ^{12}C weighing exactly 12 g. One Latin connotation for the word "mole" is "large mass" or "bulk," which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth.

The number of entities composing a mole has been experimentally determined to be $6.02214179 \times 10^{23}$, a fundamental constant named **Avogadro's number (N_A)** or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of "per mole," a conveniently rounded version being 6.022×10^{23}/mol.

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The **molar mass** of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (see **Figure 6.12**).

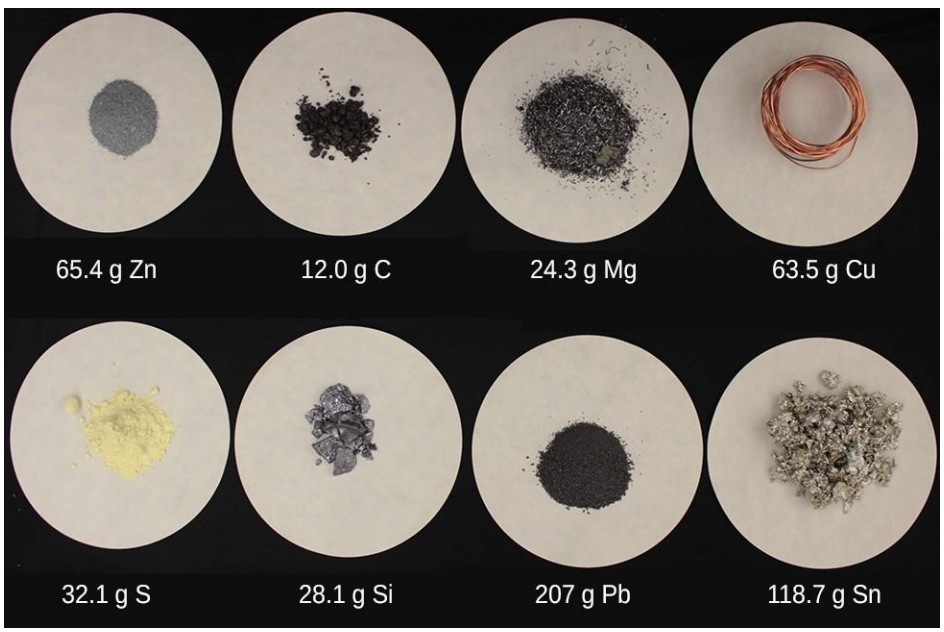


Figure 6.12 Each sample contains 6.022×10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc,12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott)

Because the definitions of both the mole and the atomic mass unit are based on the same reference substance, ^{12}C, the molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). According to the definition of the mole, 12 g of ^{12}C contains 1 mole of ^{12}C atoms (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, ^{12}C. Extending this principle, the molar mass of

a compound in grams is likewise numerically equivalent to its formula mass in amu (**Figure 6.13**).

Figure 6.13 Each sample contains 6.02×10^{23} molecules or formula units—1.00 mol of the compound or element. Clock-wise from the upper left: 130.2 g of $C_8H_{17}OH$ (1-octanol, formula mass 130.2 amu), 454.4 g of HgI_2 (mercury(II) iodide, formula mass 454.4 amu), 32.0 g of CH_3OH (methanol, formula mass 32.0 amu) and 256.5 g of S_8 (sulfur, formula mass 256.5 amu). (credit: Sahar Atwa)

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}
Cl	35.45	35.45	6.022×10^{23}

While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water weighing about 0.03 g (see **Figure 6.14**). Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules.

Figure 6.14 The number of molecules in a single droplet of water is roughly 100 billion times greater than the number of people on earth. (credit: "tanakawho"/Wikimedia commons)

Link to Learning

The mole is used in chemistry to represent 6.022×10^{23} of something, but it can be difficult to conceptualize such a large number. Watch this **video (http://openstaxcollege.org/l/16molevideo)** and then complete the "Think" questions that follow. Explore more about the mole by reviewing the information under "Dig Deeper."

The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substance's mass.

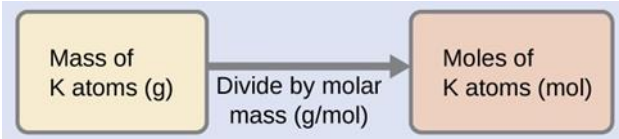
Example 6.6

Deriving Moles from Grams for an Element: According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?

Solution:

The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable "ballpark" estimate of the number of moles would be slightly greater than 0.1 mol.

The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol):



The factor-label method supports this mathematical approach since the unit "g" cancels and the answer has units of "mol:"

$$4.7g\ k\ \left(\frac{mol\ K}{39.10\ g}\right) = 0.12\ mol\ K$$

The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol.

Check Your Learning

Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g?

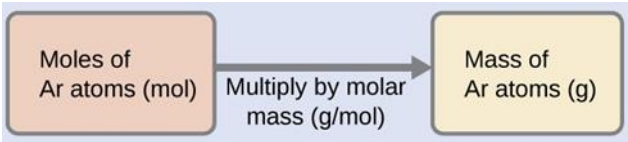
Answer: 0.360 mol

Example 6.7

Deriving Grams from Moles for an Element: A liter of air contains 9.2×10^{-4} mol argon. What is the mass of Ar in a liter of air?

Solution:

The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth ($\sim 10^{-3}$) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~0.04 g):



In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol):

$$9.2 * 10^{-4}\ mol\ Ar\ \left(\frac{39.95\ g}{mol}\right) = 0.037\ g\ Ar$$

The result is in agreement with our expectations, around 0.04 g Ar.

Check Your Learning

What is the mass of 2.561 mol of gold?

Answer: 504.4 g

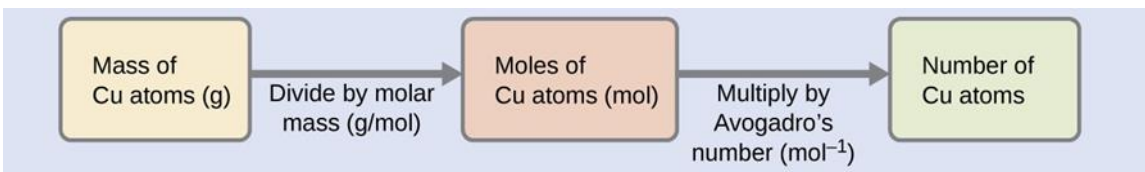
Example 6.8

Deriving Number of Atoms from Mass for an Element: Copper is commonly used to fabricate electrical wire (**Figure 6.15**). How many copper atoms are in 5.00 g of copper wire?

Figure 6.15 Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol)

Solution:

The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (N_A) to convert this molar amount to number of Cu atoms:



Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth N_A, or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$5.00\ g\ Cu\left(\frac{mol\ Cu}{63.55\ g}\right)\left(\frac{6.022 * 10^{23} atoms}{mol}\right) = 4.74 * 10^{22}\ atoms\ of\ copper$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 10^{22} as expected.

Check Your Learning

A prospector panning for gold in a river collects 15.00 g of pure gold. How many Au atoms are in this quantity of gold?

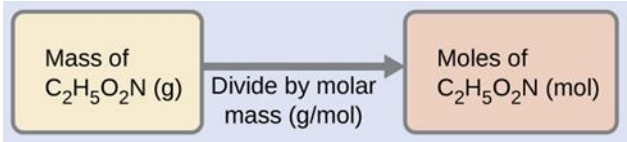
Answer: 4.586×10^{22} Au atoms

Example 6.9

Deriving Moles from Grams for a Compound: Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $C_2H_5O_2N$. How many moles of glycine molecules are contained in 28.35 g of glycine?

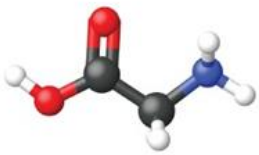
Solution:

We can derive the number of moles of a compound from its mass following the same procedure we used for an element in **Example 6.6**:



The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $C_2H_5O_2N$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07



The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$28.35\ g\ glycine\left(\frac{mol\ glycine}{75.07g}\right) = 0.378\ mol\ glycine$$

This result is consistent with our rough estimate.

Check Your Learning

How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a 25-g sample of sucrose?

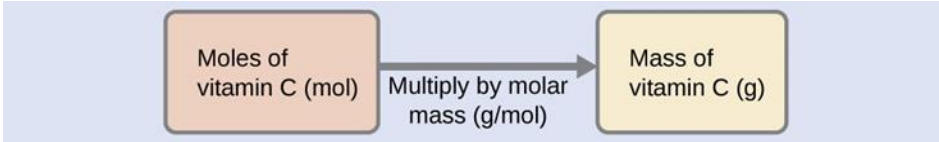
Answer: 0.073 mol

Example 6.10

Deriving Grams from Moles for a Compound: Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams?

Solution:

As for elements, the mass of a compound can be derived from its molar amount as shown:



The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole (~10−4 or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~0.02 g). Performing the calculation, we get:

$$1.42 * 10^{-4}\ mol\ vitamin\ C\ \left(\frac{176.124\ g}{mol\ vitamin\ C}\right) = 0.0250\ g$$

This is consistent with the anticipated result.

Check Your Learning

What is the mass of 0.443 mol of hydrazine, N_2H_4?

Answer: 14.2g

Assessment 6.1

There are 29 mg of caffeine in a typical cola. How many moles of caffeine are there in a can? The molecular formula for caffeine is $C_8H_{10}N_4O_2$. Strategy: Use the molecular formula to calculate the molecular mass. Then use the molecular mass to determine the number of moles of caffeine.

Step 1: Calculate molecular mass.

$$molecular\ mass = \sum atomic\ masses$$

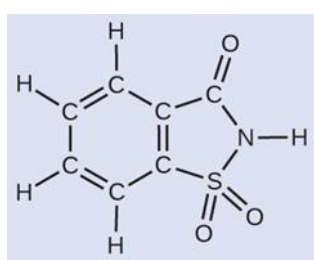
$$molecular\ mass = 8\left(12\frac{g}{mol}\right) + 10\left(1\frac{g}{mol}\right) + 4\left(14\frac{g}{mol}\right) + 2\left(16\frac{g}{mol}\right) = 194\frac{g}{mol}$$

Step 2: Use molar mass to find the number of moles of caffeine.

$$29\ mg\ caffeine \times \frac{1\ g}{1{,}000\ mg} \times \frac{1\ mol\ caffeine}{194\ g\ caffeine} = 1.49\ \times 10^{-4} moles\ of\ caffeine$$

Example 6.11

Deriving the Number of Atoms and Molecules from the Mass of a Compound: A packet of an artificial sweetener contains 40.0 mg of saccharin (C7H5NO3S), which has the structural formula:

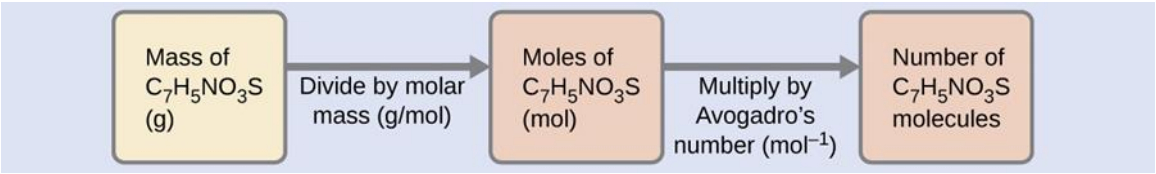


Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution:

The number of molecules in a given mass of compound is computed by first deriving the number of moles, as

demonstrated in **Example 6.9**, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400\ g\ C_7H_5NO_3S\left(\frac{mol\ C_7H_5NO_3S}{183.18\ g\ C_7H_5NO_3\ S}\right)\left(\frac{6.022*10^{23}C_7H_5NO\ S_3molecules}{1\ mol\ C_7H_5NO_3S}\right)$$
$$= 1.31*10^{20}C_7H_5NO_3S\ molecule$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number C atoms in the provided sample is:

$$1.31*10^{20}C_7H_5NO_3S\ molecules\ \left(\frac{7\ C\ atoms}{1\ C_7H_5NO_3S}\right) = 9.17*10^{20}\ C\ atoms$$

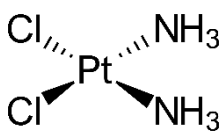
Check Your Learning

How many C_4H_{10} molecules are contained in 9.213 g of this compound? How many hydrogen atoms?

Answer: 9.545×10^{22} molecules C_4H_{10}; 9.545×10^{23} atoms H

Assessment 6.2

Cisplatin is a chemotherapy medication. Its molecular mass is 300 g/mol. If a pharmacist wants to dispense 6.00×10^{-5} moles of cisplatin, how many mg of cisplatin should she weigh out?



Strategy: use the molecular mass of cisplatin to determine the number of grams needed. Convert the grams into milligrams.

Step 1: Calculate grams of cisplatin.

$$\frac{6.00 \times 10^{-5} moles\ cisplatin}{1} \times \frac{300\ g\ cisplatin}{1\ mol\ cisplatin} = 0.0180\ g\ cisplatin$$

Step 2: Convert g to mg

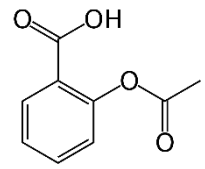
$$\frac{0.0180\ g\ cisplatin}{1} \times \frac{1{,}000\ mg}{1\ g} = 18.0\ mg\ cisplatin$$

Assessment 6.3

Aspirin, acetylsalicylic acid, is used for pain relief and after a heart attack. Its molecular formula is $C_9H_8O_4$. What is its molecular mass? How many moles of aspirin are in a typical 325 mg caplet?

Strategy: Sum the atomic masses of the constituent atoms in molecule. The atomic masses of carbon, hydrogen and oxygen are 12,1, and 16 g/mol, respectively.

Step 1: add up the atomic masses



$$Molecular\ Mass = \sum atomic\ masses$$

$$Molecular\ Mass = 9\left(12\frac{g}{mol}\right) + 8\left(1\frac{g}{mol}\right) + 4\left(16\frac{g}{mol}\right)$$

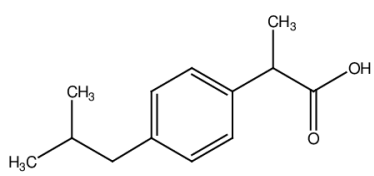
$$Molecular\ Mass\ of\ Aspirin = 180\frac{g}{mol}$$

Step 2: Calculate the number of moles of aspirin in 325 mg.

$$\frac{325\ mg\ aspirin}{1} \times \frac{1\ g}{1{,}000\ mg} \times \frac{1\ mol\ aspirin}{180\ g\ aspirin} = 0.00181\ moles\ aspirin$$

Assessment 6.4

How many molecules are in 400 mg of ibuprofen? How many **atoms** the molecular mass of ibuprofen is 206 g/mol, and its molecular formula is $C_{13}H_{18}O_2$.



Ibuprofen

Strategy: Use the molecular mass to find the number of moles, then use Avogadro's number to determine the number of molecules.

Step 1: Find the number of moles using molecular mass.

$$400\ mg\ ibuprofen \times \frac{1\ g}{1{,}000\ mg} \times \frac{1\ mole\ ibuprofen}{206\ g\ ibuprofen} = 0.00194\ moles\ of\ ibuprofen$$

Step 2: Find the number of molecules using Avogadro's number

$$0.00194\ moles\ of\ ibuprofen \times \frac{6.02 \times 10^{23}\ molecules}{1\ mol} = 1.17 \times 10^{21} molecules$$

Step 3: Use the molecular formula to find the ratio of molecules to atoms.

In each molecule, there are 13 carbons, 18 hydrogens, and 2 oxygens.

$$Number\ of\ atoms\ per\ molecule = 13 + 18 + 2 = 33\ atoms\ per\ molecule$$

Step 4: Use the number of molecules from step 2, and the ration of atoms to molecules from step 3 to find the

number of atoms.

$$1.17 \times 10^{21} molecules \times \frac{33\ atoms}{1\ molecule} = 3.86 \times 10^{22}\ atoms$$

Assessment 6.5

How many moles of heparin are there in 3.32×10^{22} molecules of heparin?

Strategy: Use Avogadro's number to find the number of moles

Solution:

$$3.32 \times 10^{22}\ molecules \times \frac{1\ mole}{6.02 \times 10^{23} molecules} = 0.0551\ moles\ of\ heparin$$

How Sciences Interconnect

Counting Neurotransmitter Molecules in the Brain

The brain is the control center of the central nervous system (**Figure 6.16**). It sends and receives signals to and from muscles and other internal organs to monitor and control their functions; it processes stimuli detected by sensory organs to guide interactions with the external world; and it houses the complex physiological processes that give rise to our intellect and emotions. The broad field of neuroscience spans all aspects of the structure and function of the central nervous system, including research on the anatomy and physiology of the brain. Great progress has been made in brain research over the past few decades, and the BRAIN Initiative, a federal initiative announced in 2013, aims to accelerate and capitalize on these advances through the concerted efforts of various industrial, academic, and government agencies (more details available at www.whitehouse.gov/share/brain-initiative).

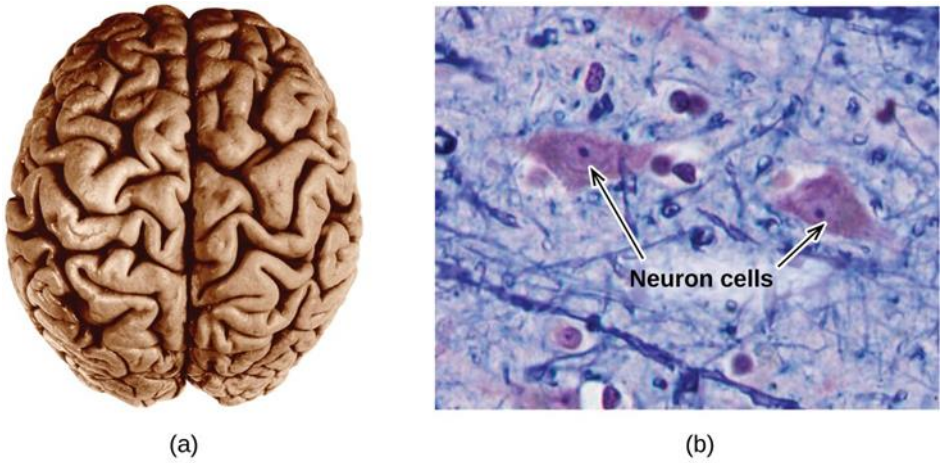


Figure 6.16 (a) A typical human brain weighs about 1.5 kg and occupies a volume of roughly 1.1 L. (b) Information is transmitted in brain tissue and throughout the central nervous system by specialized cells called neurons (micrograph shows cells at 1600× magnification).

Specialized cells called neurons transmit information between different parts of the central nervous system by way of electrical and chemical signals. Chemical signaling occurs at the interface between different neurons when one of the cells releases molecules (called neurotransmitters) that diffuse across the small gap between the cells (called the synapse) and bind to the surface of the other cell. These neurotransmitter molecules are stored in small intracellular structures called vesicles that fuse to the cell wall and then break open to release their contents when the neuron is appropriately stimulated. This process is called exocytosis (see **Figure 6.17**). One neurotransmitter that has been very extensively studied is dopamine, $C_8H_{11}NO_2$. Dopamine is involved in various neurological processes that impact a wide variety of human behaviors. Dysfunctions in the dopamine systems of the brain underlie serious neurological diseases such as Parkinson's and schizophrenia.

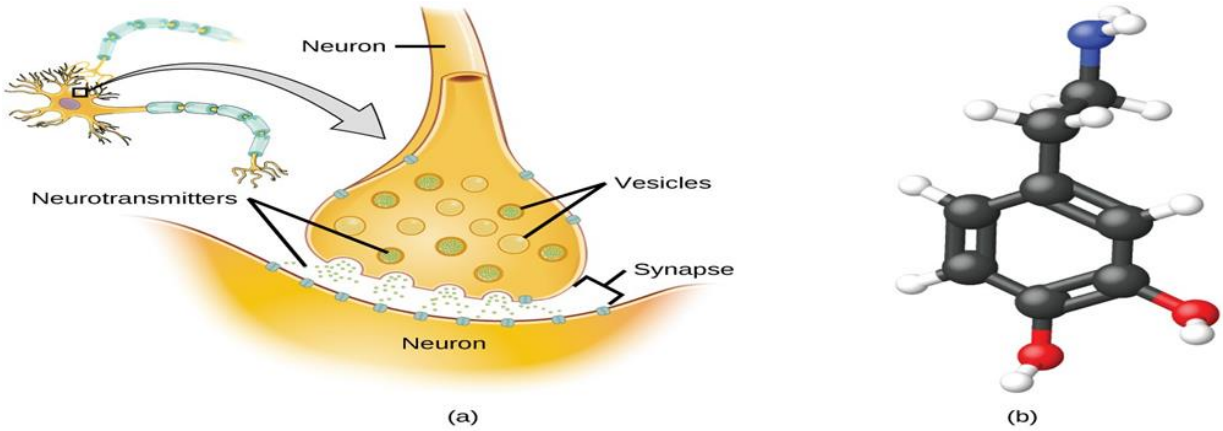


Figure 6.17 (a) Chemical signals are transmitted from neurons to other cells by the release of neurotransmitter molecules into the small gaps (synapses) between the cells. (b) Dopamine, $C_8H_{11}NO_2$, is a neurotransmitter involved in a number of neurological processes.

One important aspect of the complex processes related to dopamine signaling is the number of neurotransmitter molecules released during exocytosis. Since this number is a central factor in determining neurological response (and subsequent human thought and action), it is important to know how this number changes with certain controlled stimulations, such as the administration of drugs. It is also important to understand the mechanism responsible for any changes in the number of neurotransmitter molecules released—for example, some dysfunction in exocytosis, a change in the number of vesicles in the neuron, or a change in the number of neurotransmitter molecules in each vesicle.

Significant progress has been made recently in directly measuring the number of dopamine molecules stored in individual vesicles and the amount actually released when the vesicle undergoes exocytosis. Using miniaturized probes that can selectively detect dopamine molecules in very small amounts, scientists have determined that the vesicles of a certain type of mouse brain neuron contain an average of 30,000 dopamine molecules per vesicle (about 5×10^{-20} mol or 50 zmol). Analysis of these neurons from mice subjected to various drug therapies shows significant

changes in the average number of dopamine molecules contained in individual vesicles, increasing or decreasing by up to three-fold, depending on the specific drug used. These studies also indicate that not all of the dopamine in a given vesicle is released during exocytosis, suggesting that it may be possible to regulate the fraction released using pharmaceutical therapies.[1]

6.4 Determining Empirical and Molecular Formulas

By the end of this section, you will be able to:

- Compute the percent composition of a compound
- Determine the empirical formula of a compound
- Determine the molecular formula of a compound

In the previous section, we discussed the relationship between the bulk mass of a substance and the number of atoms or molecules it contains (moles). Given the chemical formula of the substance, we were able to determine the amount of the substance (moles) from its mass, and vice versa. But what if the chemical formula of a substance is unknown? In this section, we will explore how to apply these very same principles in order to derive the chemical formulas of unknown substances from experimental mass measurements.

Percent Composition

The elemental makeup of a compound defines its chemical identity, and chemical formulas are the most succinct way of representing this elemental makeup. When a compound's formula is unknown, measuring the mass of each of its constituent elements is often the first step in the process of determining the formula experimentally. The results of these measurements permit the calculation of the compound's **percent composition**, defined as the percentage by mass of each element in the compound. For example, consider a gaseous compound composed solely of carbon and hydrogen. The percent composition of this compound could be represented as follows:

$$\%\,H = \frac{mass\ H}{mass\ compound} * 100\%$$

$$\%\,C = \frac{mass\ C}{mass\ compound} * 100\%$$

If analysis of a 10.0-g sample of this gas showed it to contain 2.5 g H and 7.5 g C, the percent composition would be calculated to be 25% H and 75% C:

$$\%\,H = \frac{2.5\ g\ H}{10.0\ g\ compound} * 100\% = 25\%$$

$$\%\,C = \frac{7.5\ g\ C}{10.0\ g\ compound} * 10096 = 7596$$

Example 6.12

Calculation of Percent Composition:

Analysis of a 12.04-g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N. What is the percent composition of this compound?

Solution:

To calculate percent composition, we divide the experimentally derived mass of each element by the overall mass of the compound, and then convert to a percentage:

$$\%\,C = \frac{7.34\ g\ C}{12.04\ g\ compound} * 100\% =$$

$$\%\,H = \frac{1.85\ g\ H}{12.04 g\ compound} * 100\% =$$

$$\%\,N = \frac{2.85\ g\ N}{12.34 g\ compound} * 100\% =$$

The analysis results indicate that the compound is 61.0% C, 15.4% H, and 23.7% N by mass.

Check Your Learning

A 24.81-g sample of a gaseous compound containing only carbon, oxygen, and chlorine is determined to contain 3.01 g C, 4.00 g O, and 17.81 g Cl. What is this compound's percent composition?

Answer: 12.1% C, 16.1% O, 71.8% Cl

Determining Percent Composition from Formula Mass

Percent composition is also useful for evaluating the relative abundance of a given element in different compounds of known formulas. As one example, consider the common nitrogen-containing fertilizers ammonia (NH_3), ammonium nitrate (NH_4NO_3), and urea (CH_4N_2O). The element nitrogen is the active ingredient for agricultural purposes, so the mass percentage of nitrogen in the compound is a practical and economic concern for consumers choosing among these fertilizers. For these sorts of applications, the percent composition of a compound is easily derived from its formula mass and the atomic masses of its constituent elements. A molecule of NH_3 contains one N atom weighing 14.01 amu and three H atoms weighing a total of (3 × 1.008 amu) = 3.024 amu. The formula mass of ammonia is therefore (14.01 amu + 3.024 amu) = 17.03 amu, and its percent composition is:

$$\%\,N = \frac{14.01\ amu\ N}{3} * 100\% = 82.27\%$$

$$\%\,\boldsymbol{H} = \frac{\mathbf{3.024\ amu\ N}}{\mathbf{3}} * \mathbf{100}\% = \mathbf{17.76}\ \%$$

This same approach may be taken considering a pair of molecules, a dozen molecules, or a mole of molecules, etc. The latter amount is most convenient and would simply involve the use of molar masses instead of atomic and formula masses, as demonstrated **Example 6.13**. As long as we know the chemical formula of the substance in question, we can easily derive percent composition from the formula mass or molar mass.

Example 6.13

Determining Percent Composition from a Molecular Formula: Aspirin is a compound with the molecular formula C9H8O4. What is its percent composition?

Solution:

To calculate the percent composition, we need to know the masses of C, H, and O in a known mass of $C_9H_8O_4$. It is convenient to consider 1 mol of $C_9H_8O_4$ and use its molar mass (180.159 g/mole, determined from the chemical

formula) to calculate the percentages of each of its elements:

$$\%\ C = \frac{9\ mol\ C * molar\ mass\ C}{molar\ mass\ C_9H_8O_4} * 100 = \frac{9 * 12.01\frac{g}{mol}}{180.156\frac{g}{mol}} * 100 = \frac{108.09\frac{g}{mol}}{180.159\frac{g}{mol}} * 100 =$$

$$\%C = 60.00\%C$$

$$\%\ H = \frac{8\ mol\ H * molar\ mass\ H}{molar\ mass\ \ C_9H_8O_4} * 100 = \frac{8 * 1.008\frac{g}{mol}}{180.159\frac{g}{mol}} * 100 = \frac{8.\frac{064g}{mol}}{180..159\frac{g}{mol}} * 100 =$$

$$\%H = 4.476\%H$$

$$\%\ O = \frac{4\ mol\ O * molar\ mass\ O}{molar\ mass\ \ C_9H_8O_4} * 100 = \frac{4 * 16.00\frac{g}{mol}}{180.159\frac{g}{mol}} * 100 = \frac{64.00\frac{g}{mol}}{180.159\frac{g}{mol}} =$$

$$\%\ O = 35.52\%$$

Note that these percentages sum to equal 100.00% when appropriately rounded.

Check Your Learning

To three significant digits, what is the mass percentage of iron in the compound Fe_2O_3?

Answer: 69.9% Fe

Determination of Empirical Formulas

As previously mentioned, the most common approach to determining a compound's chemical formula is to first measure the masses of its constituent elements. However, we must keep in mind that chemical formulas represent the relative *numbers*, not masses, of atoms in the substance. Therefore, any experimentally derived data involving mass must be used to derive the corresponding numbers of atoms in the compound. To accomplish this, we can use molar masses to convert the mass of each element to a number of moles. We then consider the moles of each element relative to each other, converting these numbers into a whole-number ratio that can be used to derive the empirical formula of the substance. Consider a sample of compound determined to contain 1.71 g C and 0.287 g H. The corresponding numbers of atoms (in moles) are:

$$1.17\ g\ C * \frac{1\ mol\ C}{12.01\ g\ C} = 0.142\ mol$$

$$0.287\ gH * \frac{1\ mol\ H}{1.008\ g\ H} = 0.284\ mol\ H$$

Thus, we can accurately represent this compound with the formula $C_{0.142}H_{0.248}$. Of course, per accepted convention, formulas contain whole-number subscripts, which can be achieved by dividing each subscript by the smaller subscript:

$$\frac{C_{0.142}}{0.142}\quad \frac{H_{0.248}}{0.142}\ or\ CH_2$$

(Recall that subscripts of "1" are not written but rather assumed if no other number is present.)

The empirical formula for this compound is thus CH_2. This may or not be the compound's *molecular formula* as well; however, we would need additional information to make that determination (as discussed later in this section).

Consider as another example a sample of compound determined to contain 5.31 g Cl and 8.40 g O. Following the same approach yields a tentative empirical formula of:

$$Cl_{0.150}\,O_{0.525} = Cl_{\frac{0.150}{0.150}}\,O_{\frac{0.525}{0.150}} = ClO_{3.5}$$

In this case, dividing by the smallest subscript still leaves us with a decimal subscript in the empirical formula. To convert this into a whole number, we must multiply each of the subscripts by two, retaining the same atom ratio and yielding Cl_2O_7 as the final empirical formula.

In summary, empirical formulas are derived from experimentally measured element masses by:

1. Deriving the number of moles of each element from its mass
2. Dividing each element's molar amount by the smallest molar amount to yield subscripts for a tentative empirical formula
3. Multiplying all coefficients by an integer, if necessary, to ensure that the smallest whole-number ratio of subscripts is obtained

Figure 6.18 outlines this procedure in flow chart fashion for a substance containing elements A and X.

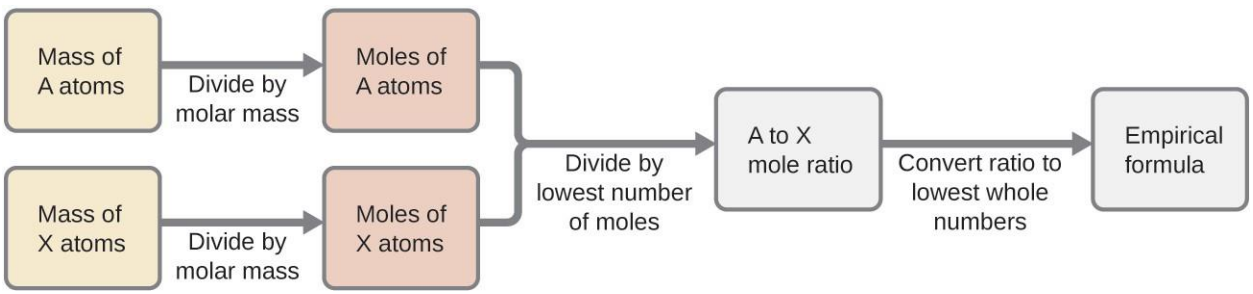


Figure 6.18 The empirical formula of a compound can be derived from the masses of all elements in the sample.

Example 6.14

Determining a Compound's Empirical Formula from the Masses of Its Elements: A sample of the black mineral hematite (**Figure 6.19**), an oxide of iron found in many iron ores, contains 34.97 g of iron and 15.03 g of oxygen. What is the empirical formula of hematite?

Figure 6.19 Hematite is an iron oxide that is used in jewelry. (credit: Mauro Cateb)

Solution:

For this problem, we are given the mass in grams of each element. Begin by finding the moles of each:

$$34.97\ g\ Fe\ \left(\frac{mol\ Fe}{55.85\ g}\right) = 0.6261\ mol\ Fe$$

$$15.03\ g\ O\ \left(\frac{mol\ O}{16.00\ g}\right) = 0.9394\ mol\ O$$

Next, derive the iron-to-oxygen molar ratio by dividing by the lesser number of moles:

$$\frac{0.6261}{0.6261} = 1.000\ mol\ Fe$$

$$\frac{0.9394}{0.6261} = 1.500\ mol\ O$$

The ratio is 1.000 mol of iron to 1.500 mol of oxygen ($Fe_1O_{1.5}$). Finally, multiply the ratio by two to get the smallest possible whole number subscripts while still maintaining the correct iron-to-oxygen ratio:

$$2\ (Fe_1O_{1.5}) = Fe_2O_3$$

The empirical formula is Fe_2O_3.

Check Your Learning

What is the empirical formula of a compound if a sample contains 0.130 g of nitrogen and 0.370 g of oxygen?

Answer: N_2O_5

Link to Learning

For additional worked examples illustrating the derivation of empirical formulas,watch the brief **video (http://openstaxcollege.org/l/16empforms)** clip.

Deriving Empirical Formulas from Percent Composition

Finally, with regard to deriving empirical formulas, consider instances in which a compound's percent composition is available rather than the absolute masses of the compound's constituent elements. In such cases, the percent composition can be used to calculate the masses of elements present in any convenient mass of compound; these

masses can then be used to derive the empirical formula in the usual fashion.

Example 6.15

Determining an Empirical Formula from Percent Composition: The bacterial fermentation of grain to produce ethanol forms a gas with a percent composition of 27.29% C and 72.71% O (**Figure 6.20**). What is the empirical formula for this gas?

Figure 6.20 An oxide of carbon is removed from these fermentation tanks through the large copper pipes at the top. (credit: "Dual Freq"/Wikimedia Commons)

Solution:

Since the scale for percentages is 100, it is most convenient to calculate the mass of elements present in a sample weighing 100 g. The calculation is "most convenient" because, per the definition for percent composition, the mass of a given element in grams is numerically equivalent to the element's mass percentage. This numerical equivalence results from the definition of the "percentage" unit, whose name is derived from the Latin phrase *per centum* meaning "by the hundred." Considering this definition, the mass percentages provided may be more conveniently expressed as fractions:

$$27.29\%\ C = \frac{27.29 g\ C}{100\ g\ compound}$$

$$72.71\ \%\ O = \frac{72.71\ g\ O}{100\ g\ compound}$$

The molar amounts of carbon and hydrogen in a 100-g sample are calculated by dividing each element's mass by its molar mass:

$$27.29\ g\ C\ \left(\frac{mol\ C}{12.01\ g}\right) = 2.272\ mol\ C$$

$$72.71\ g\ O\ \left(\frac{mol\ O}{16.00\ g}\right) = 4.544\ mol\ O$$

Coefficients for the tentative empirical formula are derived by dividing each molar amount by the lesser of the two:

$$\frac{2.272\ mol\ C}{2.272} = 1$$

$$\frac{4.544\ mol\ O}{2.272} = 2$$

Since the resulting ratio is one carbon to two oxygen atoms, the empirical formula is CO_2.

Check Your Learning

What is the empirical formula of a compound containing 40.0% C, 6.71% H, and 53.28% O?

Answer: CH_2O

Derivation of Molecular Formulas

Recall that empirical formulas are symbols representing the *relative* numbers of a compound's elements. Determining the *absolute* numbers of atoms that compose a single molecule of a covalent compound requires knowledge of both its empirical formula and its molecular mass or molar mass. These quantities may be determined experimentally by various measurement techniques. Molecular mass, for example, is often derived from the mass spectrum of the compound (see discussion of this technique in the previous chapter on atoms and molecules). Molar mass can be measured by a number of experimental methods, many of which will be introduced in later chapters of this text.

Molecular formulas are derived by comparing the compound's molecular or molar mass to its **empirical formula mass**. As the name suggests, an empirical formula mass is the sum of the average atomic masses of all the atoms represented in an empirical formula. If we know the molecular (or molar) mass of the substance, we can divide this by the empirical formula mass in order to identify the number of empirical formula units per molecule, which we designate as *n*:

$$\frac{molecular\ or\ molar\ mass\ \left(amu\ or\ \frac{g}{mol}\right)}{empirical\ formula\ mass\ \left(amu\ or\ \frac{g}{mol}\right)} = n\ formula\ \frac{units}{molecule}$$

The molecular formula is then obtained by multiplying each subscript in the empirical formula by *n*, as shown by the generic empirical formula A_xB_y:

$$(\mathrm{A_x\,B_y})_{\mathrm{n}} = \mathrm{A_{nx}\,B_{ny}}$$

For example, consider a covalent compound whose empirical formula is determined to be CH_2O. The empirical formula mass for this compound is approximately 30 amu (the sum of 12 amu for one C atom, 2 amu for two H atoms, and 16 amu for one O atom). If the compound's molecular mass is determined to be 180 amu, this indicates that molecules of this compound contain six times the number of atoms represented in the empirical formula:

$$\frac{180\,\frac{amu * molecule}{formula\ unit}}{30} = 6\ \frac{Formula\ units}{molecule}$$

Molecules of this compound are then represented by molecular formulas whose subscripts are six times greater than those in the empirical formula:

$$(\mathrm{CH_2O})_6 = \mathrm{C_6\,H_{12}\,O_6}$$

Note that this same approach may be used when the molar mass (g/mol) instead of the molecular mass (amu) is used. In this case, we are merely considering one mole of empirical formula units and molecules, as opposed to single units and molecules.

Example 6.16

Determination of the Molecular Formula for Nicotine: Nicotine, an alkaloid in the nightshade family of plants that is mainly responsible for the addictive nature of cigarettes, contains 74.02% C, 8.710% H, and 17.27% N. If 40.57 g of nicotine contains 0.2500 mol nicotine, what is the molecular formula?

Solution:

Determining the molecular formula from the provided data will require comparison of the compound's empirical formula mass to its molar mass. As the first step, use the percent composition to derive the compound's empirical formula. Assuming a convenient, a 100-g sample of nicotine yields the following molar amounts of its elements:

$$74.02\ g\ C\ \left(\frac{1\ mol\ C}{12.01\ g\ C}\right) = 6.163\ mol\ C$$

$$8.710\ g\ H\ \left(\frac{1\ mol\ H}{1.01\ g\ H}\right) = 8.624\ mol\ H$$

$$17.27\ g\ N\ \left(\frac{1\ mol\ N}{14.01\ g\ N}\right) = 1.233\ mol\ N$$

Next, we calculate the molar ratios of these elements relative to the least abundant element, N.

$$\frac{6.163\ mol\ C}{1.233\ mol\ N} = 5$$

$$\frac{8.264\ mol\ H}{1.233\ mol\ N} = 7$$

$$\frac{1.233\ mol\ N}{1.233\ mol\ N} = 1$$

The C-to-N and H-to-N molar ratios are adequately close to whole numbers, and so the empirical formula is C_5H_7N. The empirical formula mass for this compound is therefore 81.13 amu /formula unit, or 81.13 g/ mol formula unit. We calculate the molar mass for nicotine from the given mass and molar amount of compound:

$$\frac{40.57\ g\ nicotine}{0.2500\ mol\ nicotine} = \frac{162.3\ g}{mol}$$

Comparing the molar mass and empirical formula mass indicates that each nicotine molecule contains two formula units:

$$\frac{162.3\ \frac{g}{mol}}{81.13 formula\ unit} = 2\ \frac{formula\ units}{molecule}$$

Thus, we can derive the molecular formula for nicotine from the empirical formula by multiplying each subscript by two:

$$(C_5\ H_7\ N)_2 = C_{10}\ H_{14}\ N_2$$

Check Your Learning

What is the molecular formula of a compound with a percent composition of 49.47% C, 5.201% H, 28.84% N, and 16.48% O, and a molecular mass of 194.2 amu?

Answer: $C_8H_{10}N_4O_2$

6.5 Reaction Stoichiometry

By the end of this section, you will be able to:

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass, moles, and solution molarity

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's **stoichiometry**, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Food preparation, for example, offers an appropriate comparison. A recipe for making eightpancakes calls for 1 cup pancake mix, $\frac{3}{4}$ cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is:

$$1 \text{ cup mix} + \frac{3}{4}\text{cup milk} + 1 \text{ egg} \longrightarrow 8 \text{ pancakes}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24\ pancakes * \frac{1\ egg}{8\ pancakes} = 3\ eggs$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive **stoichiometric factors** that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

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

This equation shows ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2\ NH_3\ molecules}{3\ H_2 molecules}\ or\ \frac{2\ doz\ NH_3\ molecules}{3\ doz\ H_2\ molecules}\ or\ \frac{2\ mol\ NH_3 molecules}{3\ mol\ H_2 molecules}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Example 6.17

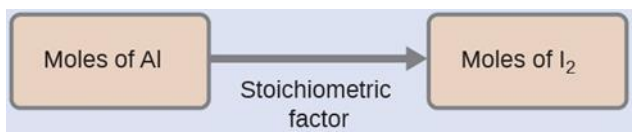
Moles of Reactant Required in a Reaction: How many moles of I_2 are required to react with 0.429 mol of Al according to the following equation (see **Figure 6.21**)?

$$2Al + 3I_2 \longrightarrow 2AlI_3$$

Figure 6.21 Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapor. (credit: modification of work by Mark Ott)

Solution:

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest Is$\frac{3\ mol\ I_2}{2\ mol\ Al}$. The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$mol\ I_2 = 0.429\ mol\ Al * \frac{3\ mol\ I_2}{2\ mol\ Al} = 0.644\ mol\ I_2$$

Check Your Learning

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation $3Ca(OH)_2 + 2H_3PO_4 \longrightarrow Ca_3(PO_4)_2 + 6H_2O$?

Answer: 2.04 mol

Example 6.18

Number of Product Molecules Generated by a Reaction: How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

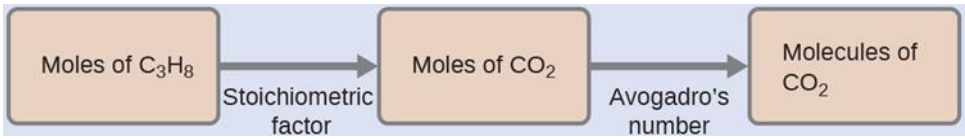
Solution:

The approach here is the same as for **Example 6.17**, though the absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

$$\frac{3\ mol\ CO_2}{1\ mol\ C_3H_8}$$

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number,



Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth NA, or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$0.75\ mol C_3H_8 * \frac{3\ mol\ CO_2}{1\ mol\ C_3H_8} * \frac{6.022 * 10^{23} molecules}{mol\ Co_2} = 1.4 * 10^{14}\ CO_2$$

Check Your Learning

How many NH_3 molecules are produced by the reaction of 4.0 mol of Ca $(OH)_2$ according to the following equation:

$$(NH_4)_2\ SO_4 + Ca\ (OH)_2 \longrightarrow 2NH_3 + CaSO_4 + 2H_2O$$

Answer: $4.8 * 10^{24}$ NH_3 molecules

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related. Directly measuring numbers of atoms and molecules is, however, not an easy task, and the practical application of stoichiometry requires that we use the more readily measured property of mass.

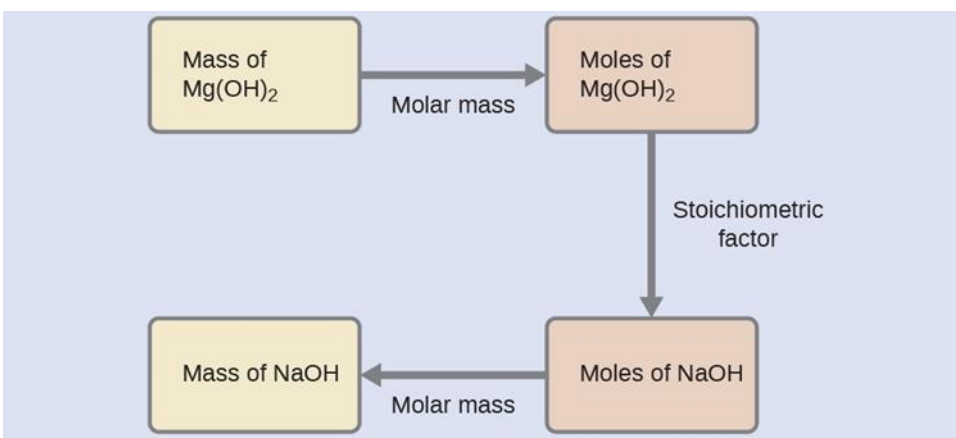
Example 6.19

Relating Masses of Reactants and Products: What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, Mg $(OH)_2$] by the following reaction?

$$MgCl_2(aq) + 2NaOH(aq) \longrightarrow Mg\ (OH)_2(s) + 2NaCl(aq)$$

Solution:

The approach used previously in Example 4.8 and Example 4.9 is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16\ g\ Mg\ (OH)_2 * \frac{1\ mol\ Mg(OH)_2}{58.3\ g\ Mg\ (OH)_2} * \frac{2\ mol\ NaOH}{1\ mol\ Mg(OH)_2} * \frac{40.0\ g\ NaOH}{mol\ NaOH} = 22g\ NaOH$$

Check Your Learning

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is $4Ga + 3O_2 \longrightarrow 2Ga_2O_3$.

Answer: 39.0 g

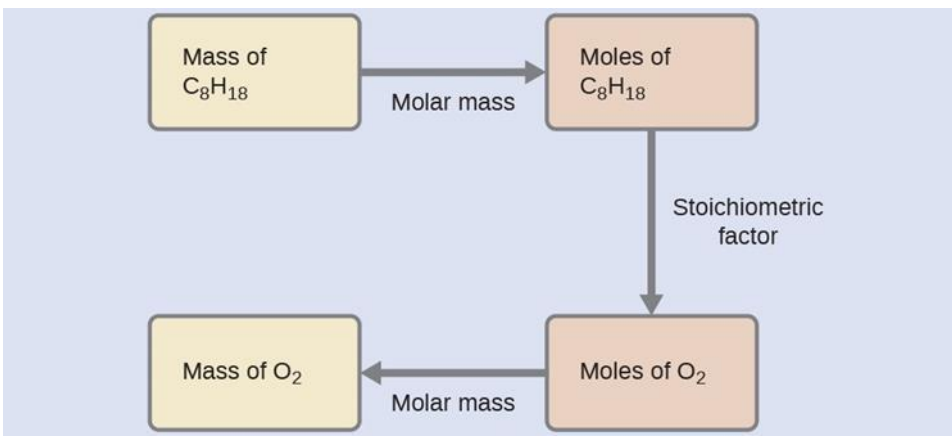
Example 6.20

Relating Masses of Reactants: What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C8H18, one of the principal components of gasoline?

$$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$$

Solution:

The approach required here is the same as for the Example 4.10, differing only in that the provided and requested masses are both for reactant species.



$$702\ g\ C_8H_{18} * \frac{1\ mol\ C_8H_{18}}{114.23\ g\ C_8H_{18}} * \frac{25\ mol\ O_2}{2\ mol\ C_8H_{18}} * \frac{32.00\ gO_2}{mol\ O_2} = 2.46 * 10^3\ gO_2$$

Check Your Learning

What mass of CO is required to react with 25.13 g of Fe2O3 according to the equation $Fe_2O_3 + 3CO \longrightarrow 2Fe^{+}$ $3CO_2$?

Answer: 13.22 g

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. **Figure 6.22** provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

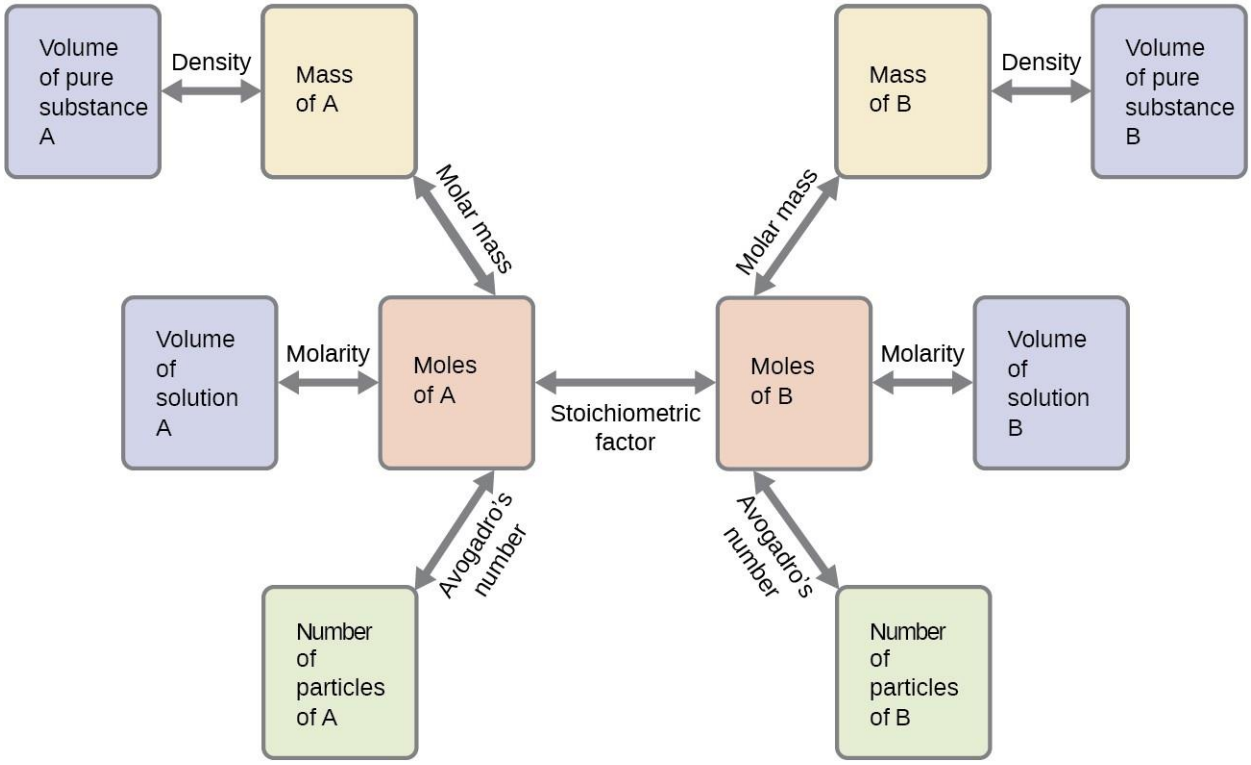


Figure 6.22 The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations.

Chemistry in everyday life

Airbags

Airbags (**Figure 6.23**) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

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

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.

Figure 6.23 Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

6.6 Reaction Yields

By the end of this section, you will be able to:

- Explain the concepts of theoretical yield and limiting reactants/reagents.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as *stoichiometric amounts*. All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Limiting Reactant

Consider another food analogy, making grilled cheese sandwiches (**Figure 6.24**):

$$1 \text{ slice of cheese} + 2 \text{ slices of bread} \longrightarrow 1 \text{ sandwich}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

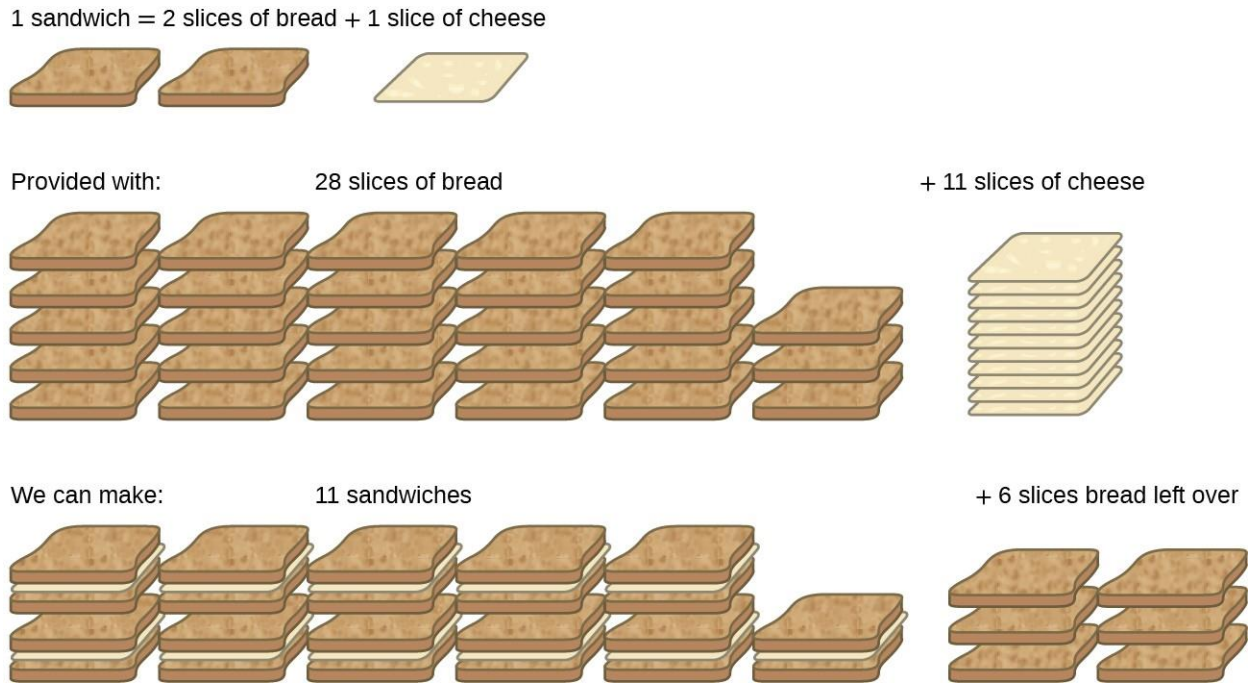


Figure 6.24 Sandwich making can illustrate the concepts of limiting and excess reactants.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

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

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the **limiting reactant**, and the other substance is the **excess reactant**. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen unreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield.

$$mol\ HCl\ produced = 3\ mol\ H_2 * \frac{2\ mol\ HCl}{1\ mol\ H_2} = 6mol\ HCl$$

Complete reaction of the provided chlorine would produce

$$mol\ HCl\ produced = 2\ mol\ Cl_2 * \frac{2\ mol\ HCl}{1\ mol\ Cl_2} = 4\ mol\ HCl$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be unreacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (**Figure 6.25**).

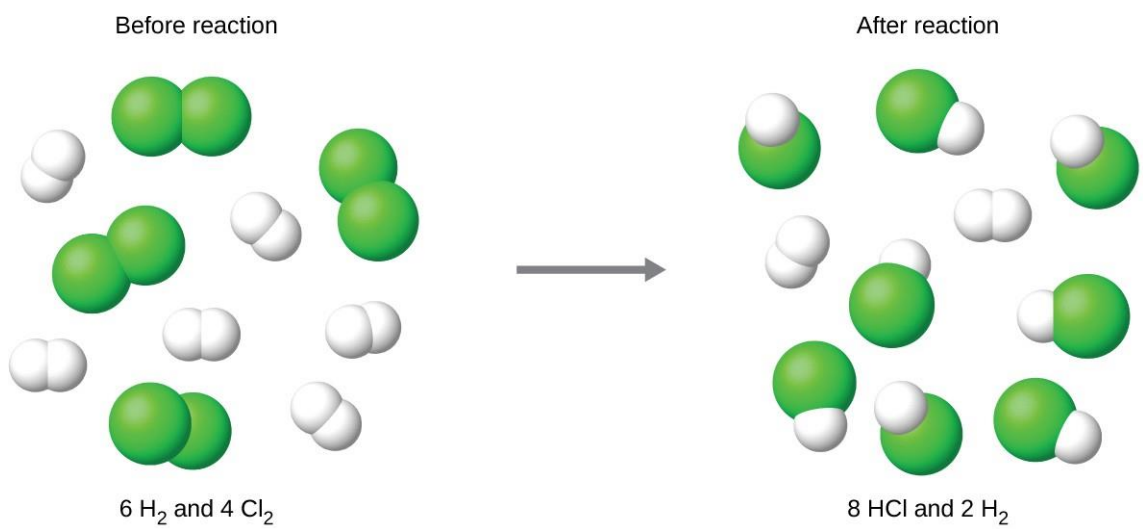


Figure 6.25 When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

Link to Learning

View this interactive **simulation (http://openstaxcollege.org/l/16reactantprod)** illustrating the concepts of limiting and excess reactants.

Example 6.21

Identifying the Limiting Reactant: Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$3Si(s) + 2N_2(g) \longrightarrow Si_3N_4(s)$

Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

Solution:

Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$mol\ Si = 2.00\ g\ Si * \frac{1\ mol\ Si}{28.09\ g\ Si} = 0.0712\ mol\ Si$$

$$mol\ N_2 = 1.50 g\ N_2 * \frac{1\ mol\ N_2}{28.02\ g\ N_2} = 0.0535\ mol\ N_2$$

The provided Si: N_2 molar ratio is:

$$\frac{0.0712\ mol\ Si}{0.0535\ mol\ N_2} = \frac{1.33\ mol\ Si}{1\ mol\ N_2}$$

The stoichiometric Si: N_2 ratio is:

$$\frac{3\ mol\ Si}{2\ mol\ N_2} = \frac{1.5\ mol\ Si}{1\ mol\ N_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$mol\ Si_3N_4\ produced = 0.0712\ mol\ Si * \frac{1\ mol\ Si_3N_4}{3\ mol\ Si} = 0.0237\ mol\ Si_3N_4$$

while the 0.0535 moles of nitrogen would produce

$$mol\ Si_3N_4\ produced = 0.0535\ mol\ N_2 * \frac{1\ mol\ Si_3\ N_4}{2\ mol\ Si} = 0.0268\ mol\ Si_3N_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant.

Check Your Learning

Which is the limiting reactant when 5.00 g of H_2 and 10.0 g of O_2 react and form water?

Answer: O_2

Percent Yield

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the **theoretical yield** of the reaction. In practice, the amount of product obtained is called the **actual yield**, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this chapter). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its **percent yield**:

$$Percent\ Yield = \frac{actual\ yield}{theoretical\ yield} * 100\%$$

Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when percent yield is calculated.

Example 6.22

Calculation of Percent Yield: Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$CuSO_4(aq) + Zn(s) \longrightarrow Cu(s) + ZnSO_4(aq)$$

What is the percent yield?

Solution:

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

$$1.274\ g\ CuSO_4 * \frac{1\ mol\ CuSO_4}{159.62\ g\ CuSO_4} * \frac{1\ mol\ Cu}{1\ mol\ CuSO_4} * \frac{63.55\ g\ Cu}{1\ mol\ Cu} = 0.5072\ g\ Cu$$

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be

$$percent\ yield = \left(\frac{actual\ yield}{theoretical\ yield}\right) * 100$$

$$Percent\ yield = \left(\frac{0.392\ g\ Cu}{0.5072\ g\ cu}\right) * 100 = 77.3\%$$

Check Your Learning

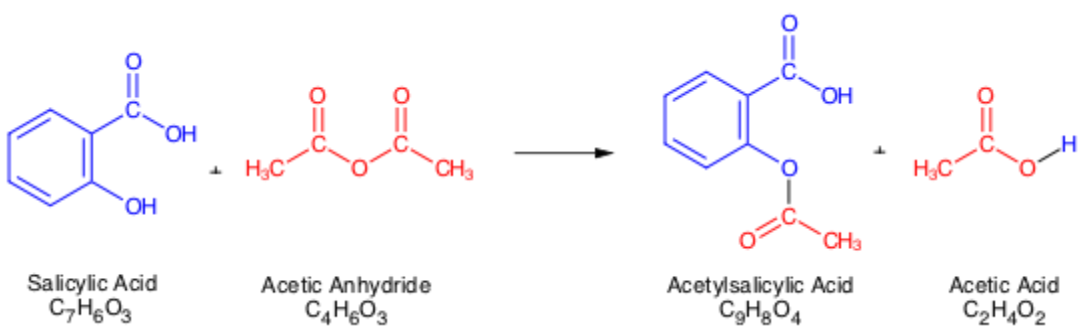
What is the percent yield of a reaction that produces 12.5 g of the gas Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$CCl_4 + 2HF \longrightarrow CF_2Cl_2 + 2HCl$

Answer: 48.3%

Assessment 6.6

Aspirin is made by reacting salicylic acid ($C_7H_6O_3$) with acetic anhydride ($C_4H_6O_3$) to form acetylsalicylic acid (aspirin $C_9H_8O_4$) and acetic acid ($C_2H_4O_2$). I 1.74 grams of acetic anhydride are allowed to react with 1.92 g of salicylic acid, what is the limiting reactant? How much excess reactant remains? If 2.14 grams of aspirin are recovered, what is the percent yield for the reaction?



Strategy: Use molecular masses to fine the moles of the salicylic acid and acetic anhydride. Use the balanced chemical equation to find the smallest number of moles produced of product from the moles of starting materials. Use the mole to mole ratio between the limited and excess reagents and the number of moles of the limiting reagent to find the moles of excess reactant consumed. Subtract the moles of excess reactant consumed from the initial moles of that reactant to find the moles remaining. Use the moles of product formed to find grams of aspirin

produced, and the actual grams of aspirin produced to find percent yield.

Step 1: Find the initial moles of the reactants

$$1.92\ g\ C_7H_6O_3 \times \frac{1\ mole\ C_7H_6O_3}{138\ g\ C_7H_6O_3} = 0.0139\ mol\ C_7H_6O_3\ (salicylic\ acid)$$

$$1.74\ g\ C_7H_6O_3 \times \frac{1\ mole\ C_4H_6O_3}{104\ g\ C_4H_6O_3} = 0.0167\ mol\ C_4H_6O_3\ (acetic\ anhydride)$$

Step 2: Determine the limiting reactant using moles of reactants, and the balanced chemical equation.

$$0.0139\ mol\ C_7H_6O_3\ (salicylic\ acid) \times \frac{1\ mole\ C_9H_8O_4}{1\ mol\ C_7H_6O_3} = 0.0139\ moles\ C_9H_8O_4\ (aspirin)$$

$$0.0167\ mol\ C_4H_6O_3\ (acetic\ anhydride) \times \frac{1\ mol\ C_9H_8O_4}{1\ mol\ C_4H_6O_3} = 0.0167\ mol\ C_9H_8O_4\ (aspirin)$$

$$0.0139\ moles\ C_9H_8O_4\ (\ from\ salicylic\ acid) < 0.0167\ mol\ C_9H_8O_4\ \ (from\ acetic\ anydride)$$

Therefore, salicylic acid is the limiting reactant.

Step 3: Find moles of excess reactant consumed.

$$0.0139\ moles\ C_9H_8O_4 \times \frac{1\ mol\ acetic\ anhydride}{1\ mol\ salicylic\ acid} = 0.0139\ moles\ of\ acetic\ anhydride\ consumed$$

Step 4: Use subtraction to find the moles of excess reactant remaining, then convert to grams.

$$0.0167\ mol\ C_4H_6O_3(initial) - \ 0.0139\ moles\ C_4H_6O_3(used) = 0.00283\ moles\ C_4H_6O_3\ (left\ over)$$

Convert to grams

$$0.00283\ moles\ C_4H_6O_3 \times \frac{104\ g\ C_4H_6O_3}{1\ mole\ \ C_4H_6O_3} = 0.294\ g\ C_4H_6O_3\ (acetic\ anhydride)\ remain.$$

Step 5: Determine the mass of aspirin formed (i.e. the *theoretical yield*)

$$0.0139\ moles\ C_9H_8O_4 \times \frac{180\ g\ C_9H_8O_4}{1\ mole\ C_9H_8O_4} = 2.50\ g\ C_9H_8O_4\ (aspirin, theoretical\ yield)$$

Step 6: Calculate percent yield.

$$\%Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100$$

$$\%Yield = \frac{2.14\ g\ aspirin}{2.50\ g\ aspirin} \times 100 = 85.5\%\ yield$$

How Sciences Interconnect

Green Chemistry and Atom Economy

The purposeful design of chemical products and processes that minimize the use of environmentally hazardous substances and the generation of waste is known as *green chemistry*. Green chemistry is a philosophical approach that is being applied to many areas of science and technology, and its practice is summarized by guidelines known as the "Twelve Principles of Green Chemistry" (see details at this **website (http://openstaxcollege.org/l/16greenchem)**). One of the 12 principles is aimed specifically at maximizing the efficiency of processes for synthesizing chemical products. The *atom economy* of a process is a measure of this

efficiency, defined as the percentage by mass of the final product of a synthesis relative to the masses of *all* the reactants used:

$$atom\ economy = \frac{mass\ of\ product}{mass\ reactants} * 100\%$$

Though the definition of atom economy at first glance appears very similar to that for percent yield, be aware that this property represents a difference in the *theoretical* efficiencies of *different* chemical processes. The percent yield of a given chemical process, on the other hand, evaluates the efficiency of a process by comparing the yield of product actually obtained to the maximum yield predicted by stoichiometry.

The synthesis of the common nonprescription pain medication, ibuprofen, nicely illustrates the success of a green chemistry approach (**Figure 6.26**). First marketed in the early 1960s, ibuprofen was produced using a six-step synthesis that required 514 g of reactants to generate each mole (206 g) of ibuprofen, an atom economy of 40%. In the 1990s, an alternative process was developed by the BHC Company (now BASF Corporation) that requires only three steps and has an atom economy of ~80%, nearly twice that of the original process. The BHC process generates significantly less chemical waste; uses less-hazardous and recyclable materials; and provides significant cost-savings to the manufacturer (and, subsequently, the consumer). In recognition of the positive environmental impact of the BHC process, the company received the Environmental Protection Agency's Greener Synthetic Pathways Award in 1997.

(a) (b)

Figure 6.26 (a) Ibuprofen is a popular nonprescription pain medication commonly sold as 200 mg tablets. (b) The BHC process for synthesizing ibuprofen requires only three steps and exhibits an impressive atom economy. (credit a: modification of work by Derrick Coetzee)

Key Terms

actual yield amount of product formed in a reaction

Avogadro's number (N_A) experimentally determined value of the number of entities comprising 1 mole of substance, equal to 6.022×10^{23} mol^{-1}

balanced equation chemical equation with equal numbers of atoms for each element in the reactant and product

chemical equation symbolic representation of a chemical reaction

coefficient number placed in front of symbols or formulas in a chemical equation to indicate their relative amount

combustion reaction vigorous redox reaction producing significant amounts of energy in the form of heat and, sometimes, light

empirical formula mass sum of average atomic masses for all atoms represented in an empirical formula

excess reactant reactant present in an amount greater than required by the reaction stoichiometry

formula mass sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass

half-reaction an equation that shows whether each reactant loses or gains electrons in a reaction.

limiting reactant reactant present in an amount lower than required by the reaction stoichiometry, thus limiting the amount of product generated

molar mass mass in grams of 1 mole of a substance

mole amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

oxidation number (also, oxidation state) the charge each atom of an element would have in a compound if the compound were ionic

oxidation process in which an element's oxidation number is increased by loss of electrons

oxidation-reduction reaction (also, redox reaction) reaction involving a change in oxidation number for one or more reactant elements

oxidizing agent (also, oxidant) substance that brings about the oxidation of another substance, and in the process becomes reduced

percent yield measure of the efficiency of a reaction, expressed as a percentage of the theoretical yield

product substance formed by a chemical or physical change; shown on the right side of the arrow in a chemical equation

quantitative analysis the determination of the amount or concentration of a substance in a sample

reactant substance undergoing a chemical or physical change; shown on the left side of the arrow in a chemical equation

reducing agent (also, reductant) substance that brings about the reduction of another substance, and in the process becomes oxidized

reduction process in which an element's oxidation number is decreased by gain of electrons

single-displacement reaction (also, replacement) redox reaction involving the oxidation of an elemental substance by an ionic species

stoichiometric factor ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products

stoichiometry relationships between the amounts of reactants and products of a chemical reaction

theoretical yield amount of product that may be produced from a given amount of reactant(s) according to the reaction stoichiometry

Key Equations

- $\boldsymbol{percent\ yield = \left(\frac{actual\ yield}{theoretical}\right) * 100}$
- $\boldsymbol{\frac{molecular\ or\ molar\ mass\ \left(amu\ or \frac{g}{mol}\right)}{emperical\ formula\ mass\ \left(amu\ or \frac{g}{mol}\right)} = n\ \frac{formula\ units}{molecules}}$

Summary

6.1 Writing and Balancing Chemical Equations

Chemical equations are symbolic representations of chemical and physical changes. Formulas for the substances undergoing the change (reactants) and substances generated by the change (products) are separated by an arrow and preceded by integer coefficients indicating their relative numbers. Balanced equations are those whose coefficients result in equal numbers of atoms for each element in the reactants and products. Chemical reactions in aqueous solution that involve ionic reactants or products may be represented more realistically by complete ionic equations and, more succinctly, by net ionic equations.

6.2 Classifying Chemical Reactions

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

6.3 Formula Mass and the Mole Concept

The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. A convenient amount unit for expressing very large numbers of atoms or molecules is the mole. Experimental

measurements have determined the number of entities composing 1 mole of substance to be 6.022×10^{23}, a quantity called Avogadro's number. The mass in grams of 1 mole of substance is its molar mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H_2O molecule weighs approximately18 amu and 1 mole of H_2O molecules weighs approximately 18 g).

6.4 Determining Empirical and Molecular Formulas

The chemical identity of a substance is defined by the types and relative numbers of atoms composing its fundamental entities (molecules in the case of covalent compounds, ions in the case of ionic compounds). A compound's percent composition provides the mass percentage of each element in the compound, and it is often experimentally determined and used to derive the compound's empirical formula. The empirical formula mass of a covalent compound may be compared to the compound's molecular or molar mass to derive a molecular formula.

6.5 Reaction Stoichiometry

A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties.

6.6 Reaction Yields

When reactions are carried out using less-than-stoichiometric quantities of reactants, the amount of product generated will be determined by the limiting reactant. The amount of product generated by a chemical reaction is its actual yield. This yield is often less than the amount of product predicted by the stoichiometry of the balanced chemical equation representing the reaction (its theoretical yield). The extent to which a reaction generates the theoretical amount of product is expressed as its percent yield.

Exercises

6.1 Writing and Balancing Chemical Equations

(a) What does it mean to say an equation is balanced? Why is it important for an equation to be balanced?

(b) Consider molecular, complete ionic, and net ionic equations.

(c) What is the difference between these types of equations?

(d) In what circumstance would the complete and net ionic equations for a reaction be identical?

2 Balance the following equations:

(a) $PCl_5(s) + H_2O(l) \longrightarrow POCl_3(l) + HCl(aq)$

(b) $Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(l) + NO(g)$

(c) $H_2(g) + I_2(s) \longrightarrow HI(s)$

(d) $Fe(s) + O_2(g) \longrightarrow Fe_2O_3(s)$

(e) $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$

(f) $(NH_4)_2Cr_2O_7(s) \longrightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$

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

(h) $PtCl_4(s) \longrightarrow Pt(s)+Cl_2(g)$

3 Balance the following equations:

(a) $Ag(s) + H_2S(g) + O_2(g) \longrightarrow Ag_2S(s) + H_2O(l)$

(b) $P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$

(c) $Pb(s) + H_2O(l) + O_2(g) \longrightarrow Pb(OH)_2(s)$

(d) $Fe(s) + H_2O(l) \longrightarrow Fe_3O_4(s) + H_2(g)$

(e) $Sc_2O_3(s) + SO_3(l) \longrightarrow Sc_2(SO_4)_3(s)$

(f) $Ca_3(PO_4)_2(aq) + H_3PO_4(aq) \longrightarrow Ca(H_2PO_4)_2(aq)$

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

(h) $TiCl_4(s) + H_2O(g) \longrightarrow TiO_2(s) + HCl(g)$

4 Write a balanced molecular equation describing each of the following chemical reactions.

(a) Solid calcium carbonate is heated and decomposes to solid calcium oxide and carbon dioxide gas.

(b) Gaseous butane, C_4H_{10}, reacts with diatomic oxygen gas to yield gaseous carbon dioxide and water vapor.

(c) Aqueous solutions of magnesium chloride and sodium hydroxide react to produce solid magnesium hydroxide and aqueous sodium chloride.

(d) Water vapor reacts with sodium metal to produce solid sodium hydroxide and hydrogen gas.

5 Write a balanced equation describing each of the following chemical reactions.

(a) Solid potassium chlorate, $KClO_3$, decomposes to form solid potassium chloride and diatomic oxygen gas.

(b) Solid aluminum metal reacts with solid diatomic iodine to form solid Al_2I_6.

6 Colorful fireworks often involve the decomposition of barium nitrate and potassium chlorate and the reaction of the metals magnesium, aluminum, and iron with oxygen.

(a) Write the formulas of barium nitrate and potassium chlorate.

(b) The decomposition of solid potassium chlorate leads to the formation of solid potassium chloride and diatomic oxygen gas. Write an equation for the reaction.

(c) The decomposition of solid barium nitrate leads to the formation of solid barium oxide, diatomic nitrogen gas, and diatomic oxygen gas. Write an equation for the reaction.

(d) Write separate equations for the reactions of the solid metals magnesium, aluminum, and iron with diatomic oxygen gas to yield the corresponding metal oxides. (Assume the iron oxide contains Fe^{3+} ions.)

7 A novel process for obtaining magnesium from sea water involves several reactions. Write a balanced chemical equation for each step of the process.

(a) The first step is the decomposition of solid calcium carbonate from seashells to form solid calcium oxide and gaseous carbon dioxide.

(b) The second step is the formation of solid calcium hydroxide as the only product from the reaction of the solid calcium oxide with liquid water.

(c) Solid calcium hydroxide is then added to the seawater, reacting with dissolved magnesium chloride to yield solid magnesium hydroxide and aqueous calcium chloride.

(d) The solid magnesium hydroxide is added to a hydrochloric acid solution, producing dissolved magnesium chloride and liquid water.

(e) Finally, the magnesium chloride is melted and electrolyzed to yield liquid magnesium metal and diatomic chlorine gas.

6.2 Classifying Chemical Reactions

8 Use the following equations to answer the next four questions:

i. $H_2O(s) \longrightarrow H_2O(l)$

ii. $Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$

iii. $CH_3OH(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$

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

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

(a) Which equation describes a physical change?

(b) Which equation identifies the reactants and products of a combustion reaction?

(c) Which equation is not balanced?

9 Indicate what type, or types, of reaction each of the following represents:

(a) $Ca(s) + Br_2(l) \longrightarrow CaBr_2(s)$

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

(c) $C_6H_{12}(l) + 9O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(g)$

10 Indicate what type, or types, of reaction each of the following represents:

(a) $H_2O(g) + C(s) \longrightarrow CO(g) + H_2(g)$

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

(c) $Al(OH)_3(aq) + 3HCl(aq) \longrightarrow AlCl_3(aq) + 3H_2O(l)$

(d) $Pb(NO_3)_2(aq) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + 2HNO_3(aq)$

11 Determine the oxidation states of the elements in the following compounds:

(a) NaI

(b) $GdCl_3$

(c) $LiNO_3$

(d) H_2Se

(e) Mg_2Si

(f) RbO_2, rubidium superoxide

(g) HF

12 Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing

compounds are peroxides or superoxides.

(a) H_3PO_4

(b) $Al(OH)_3$

(c) SeO_2

(d) KNO_2

(e) In_2S_3

(f) P_4O_6

13 Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing compounds are peroxides or superoxides.

(a) H_2SO_4

(b) $Ca(OH)_2$

(c) BrOH

(d) $ClNO_2$

(e) $TiCl_4$

(f) NaH

14 Identify the atoms that are oxidized and reduced, the change in oxidation state for each, and the oxidizing and reducing agents in each of the following equations:

(a) $Mg(s) + NiCl_2(aq) \longrightarrow MgCl_2(aq) + Ni(s)$

(b) $PCl_3(l) + Cl_2(g) \longrightarrow PCl_5(s)$

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

(d) $Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$

(e) $2K_2S_2O_3(s) + I_2(s) \longrightarrow K_2S_4O_6(s) + 2KI(s)$

(f) $3Cu(s) + 8HNO_3(aq) \longrightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$

15 Complete and balance the following oxidation-reduction reactions, which give the highest possible oxidation state for the oxidized atoms.

(a) $Al(s) + F_2(g) \longrightarrow$

(b) $Al(s) + CuBr_2(aq) \longrightarrow$ (singledisplacemt)

(c) (c) $P_4(s) + O_2(g) \longrightarrow$

(d) $Ca(s) + H_2O(l) \longrightarrow$ products are a strong base and a diatomic gas)

16 Complete and balance the following oxidation-reduction reactions, which give the highest possible oxidation state for the oxidized atoms.

(a) $K(s) + H_2O(l) \longrightarrow$

(b) $Ba(s) + HBr(aq) \longrightarrow$

(c) $Sn(s) + I_2(s) \longrightarrow$

17 When heated to 700–800 °C, diamonds, which are pure carbon, are oxidized by atmospheric oxygen.

(They burn!) Write the balanced equation for this reaction.

18 The military has experimented with lasers that produce very intense light when fluorine combines explosively with hydrogen. What is the balanced equation for this reaction?

19 Great Lakes Chemical Company produces bromine, Br_2, from bromide salts such as NaBr, in Arkansas brine by treating the brine with chlorine gas. Write a balanced equation for the reaction of NaBr with Cl_2.

20 In a common experiment in the general chemistry laboratory, magnesium metal is heated in air to produce MgO. MgO is a white solid, but in these experiments, it often looks gray, due to small amounts of Mg_3N_2, a compound formed as some of the magnesium reacts with nitrogen. Write a balanced equation for each reaction.

6.3 Formula Mass and the Mole Concept

1. What is the total mass (amu) of carbon in each of the following molecules?

(a) CH_4

(b) $CHCl_3$

(c) $C_{12}H_{10}O_6$

(d) $CH_3CH_2CH_2CH_2CH_3$

2. What is the total mass of hydrogen in each of the molecules?

(a) CH_4

(b) $CHCl_3$

(c) $C_{12}H_{10}O_6$

(d) $CH_3CH_2CH_2CH_2CH_3$

3. Calculate the molecular or formula mass of each of the following:

(a) P_4

(b) H_2O

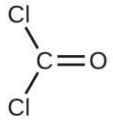
(c) $Ca(NO_3)_2$

(d) CH_3CO_2H (acetic acid)

(e) $C_{12}H_{22}O_{11}$ (sucrose, cane sugar).

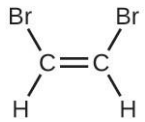
4. Determine the molecular mass of the following compounds:

(a)

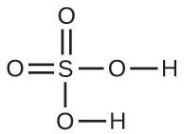


(b)

(c)

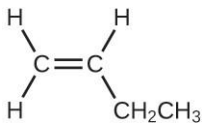


(d)

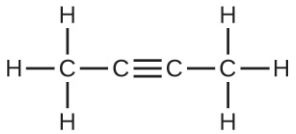


5. Determine the molecular mass of the following compounds:

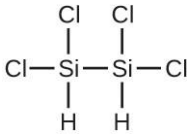
(a)



(b)



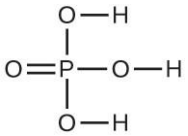
(c)



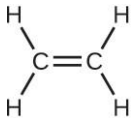
(d)

6. Which molecule has a molecular mass of 28.05 amu?

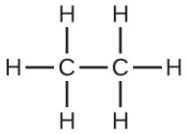
(a)



(b)



(c)



7. Write a sentence that describes how to determine the number of moles of a compound in a known mass of the compound if we know its molecular formula.

8. Compare 1 mole of H_2, 1 mole of O_2, and 1 mole of F_2.

(a) Which has the largest number of molecules? Explain why.

(b) Which has the greatest mass? Explain why.

9. Which contains the greatest mass of oxygen: 0.75 mol of ethanol (C_2H_5OH), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

10. Which contains the greatest number of moles of oxygen atoms: 1 mol of ethanol (C_2H_5OH), 1 mol of formic acid (HCO_2H), or 1 mol of water (H_2O)? Explain why.

11. How are the molecular mass and the molar mass of a compound similar and how are they different?

12. Calculate the molar mass of each of the following compounds:

(a) hydrogen fluoride, HF

(b) ammonia, NH_3

(c) nitric acid, HNO_3

(d) silver sulfate, Ag_2SO_4

(e) boric acid, $B(OH)_3$

13. Calculate the molar mass of each of the following:

(a) S_8

(b) C_5H_{12}

(c) $Sc_2(SO_4)_3$

(d) CH_3COCH_3 (acetone)

(e) $C_6H_{12}O_6$ (glucose)

14. Calculate the empirical or molecular formula mass and the molar mass of each of the following minerals:

(a) limestone, $CaCO_3$

(b) halite, NaCl

(c) beryl, $Be_3Al_2Si_6O_{18}$

(d) malachite, $Cu_2(OH)_2CO_3$

(e) turquoise, $CuAl_6(PO_4)_4(OH)_8(H_2O)_4$

15. Calculate the molar mass of each of the following:

(a) the anesthetic halothane, $C_2HBrClF_3$

(b) the herbicide paraquat, $C_{12}H_{14}N_2Cl_2$

(c) caffeine, $C_8H_{10}N_4O_2$

(d) urea, $CO(NH_2)_2$

(e) a typical soap, $C_{17}H_{35}CO_2Na$

16. Determine the number of moles of compound and the number of moles of each type of atom in each of the following:

(a) 25.0 g of propylene, C_3H_6

(b) 3.06×10^{-3} g of the amino acid glycine, $C_2H_5NO_2$

(c) 25 lb of the herbicide Treflan, $C_{13}H_{16}N_2O_4F$ (1 lb = 454 g)

(d) 0.125 kg of the insecticide Paris Green, $Cu_4(AsO_3)_2(CH_3CO_2)_2$

(e) 325 mg of aspirin, $C_6H_4(CO_2H)(CO_2CH_3)$

17. Determine the mass of each of the following:

(a) 0.0146 mol KOH

(b) 10.2 mol ethane, C_2H_6

(c) 1.6×10^{-3} mol Na_2SO_4

(d) 6.854×10^3 mol glucose, $C_6H_{12}O_6$

(e) 2.86 mol $Co(NH_3)_6Cl_3$

18. Determine the number of moles of the compound and determine the number of moles of each type of atom in each of the following:

(a) 2.12 g of potassium bromide, KBr

(b) 0.1488 g of phosphoric acid, H_3PO_4

(c) 23 kg of calcium carbonate, $CaCO_3$

(d) 78.452 g of aluminum sulfate, $Al_2(SO_4)_3$

(e) 0.1250 mg of caffeine, $C_8H_{10}N_4O_2$

19. Determine the mass of each of the following:

(a) 2.345 mol LiCl

(b) 0.0872 mol acetylene, C_2H_2

(c) 3.3×10^{-2} mol Na_2CO_3

(d) 1.23×10^3 mol fructose, $C_6H_{12}O_6$

(e) 0.5758 mol $FeSO_4(H_2O)_7$

20. The approximate minimum daily dietary requirement of the amino acid leucine, $C_6H_{13}NO_2$, is 1.1 g. What is this requirement in moles?

21. Determine the mass in grams of each of the following:

(a) 0.600 mol of oxygen atoms

(b) 0.600 mol of oxygen molecules, O_2

(c) 0.600 mol of ozone molecules, O_3

22. A 55-kg woman has 7.5×10^{-3} mol of hemoglobin (molar mass = 64,456 g/mol) in her blood. How many hemoglobin molecules is this? What is this quantity in grams?

23. Determine the number of atoms and the mass of zirconium, silicon, and oxygen found in 0.3384 mol of zircon, $ZrSiO_4$, a semiprecious stone.

24. Determine which of the following contains the greatest mass of hydrogen: 1 mol of CH_4, 0.6 mol of C_6H_6, or 0.4 mol of C_3H_8.

25. Determine which of the following contains the greatest mass of aluminum: 122 g of $AlPO_4$, 266 g of Al_2C1_6, or 225 g of Al_2S_3.

26. Diamond is one form of elemental carbon. An engagement ring contains a diamond weighing 1.25 carats (1 carat = 200 mg). How many atoms are present in the diamond?

27. The Cullinan diamond was the largest natural diamond ever found (January 25, 1905). It weighed 3104 carats (1 carat = 200 mg). How many carbon atoms were present in the stone?

28. One 55-gram serving of a particular cereal supplies 270 mg of sodium, 11% of the recommended daily allowance. How many moles and atoms of sodium are in the recommended daily allowance?

29. A certain nut crunch cereal contains 11.0 grams of sugar (sucrose, $C_{12}H_{22}O_{11}$) per serving size of 60.0 grams. How many servings of this cereal must be eaten to consume 0.0278 moles of sugar?

30. A tube of toothpaste contains 0.76 g of sodium monofluorophosphate (Na_2PO_3F) in 100 mL.

(a) What mass of fluorine atoms in mg was present?

(b) How many fluorine atoms were present?

31. Which of the following represents the least number of molecules? (a) 20.0 g of H_2O (18.02 g/mol)

(b) 77.0 g of CH_4 (16.06 g/mol)

(c) 68.0 g of CaH_2 (42.09 g/mol) (d) 100.0 g of N_2O (44.02 g/mol)

(e) 84.0 g of HF (20.01 g/mol)

6.4 Determining Empirical and Molecular Formulas

32. What information do we need to determine the molecular formula of a compound from the empirical formula?

33. Calculate the following to four significant figures:

(a) the percent composition of ammonia, NH_3

(b) the percent composition of photographic fixer solution ("hypo"), $Na_2S_2O_3$

(c) the percent of calcium ion in $Ca_3(PO_4)_2$

34. Determine the following to four significant figures:

(a) the percent composition of hydrazoic acid, HN_3

(b) the percent composition of TNT, $C_6H_2(CH_3)(NO_2)_3$

(c) the percent of SO_4^{2-} in $Al_2(SO_4)_3$

35. Determine the percent ammonia, NH_3, in $Co(NH_3)_6Cl_3$, to three significant figures.

36. Determine the percent water in $CuSO_4{\cdot}5H_2O$ to three significant figures.

37. Determine the empirical formulas for compounds with the following percent compositions:

(a) 15.8% carbon and 84.2% sulfur

(b) 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen

38. Determine the empirical formulas for compounds with the following percent compositions:

(a) 43.6% phosphorus and 56.4% oxygen

(b) 28.7% K, 1.5% H, 22.8% P, and 47.0% O

39. A compound of carbon and hydrogen contains 92.3% C and has a molar mass of 78.1 g/mol. What is its molecular formula?

40. Dichloroethane, a compound that is often used for dry cleaning, contains carbon, hydrogen, and chlorine. It has a molar mass of 99 g/mol. Analysis of a sample shows that it contains 24.3% carbon and 4.1% hydrogen. What is its molecular formula?

41. Determine the empirical and molecular formula for chrysotile asbestos. Chrysotile has the following percent composition: 28.03% Mg, 21.60% Si, 1.16% H, and 49.21% O. The molar mass for chrysotile is 520.8 g/mol.

42. Polymers are large molecules composed of simple units repeated many times. Thus, they often have relatively simple empirical formulas. Calculate the empirical formulas of the following polymers:

(a) Lucite (Plexiglas); 59.9% C, 8.06% H, 32.0% O (b) Saran; 24.8% C, 2.0% H, 73.1% Cl

(c) polyethylene; 86% C, 14% H

(d) polystyrene; 92.3% C, 7.7% H

(e) Orlon; 67.9% C, 5.70% H, 26.4% N

43. A major textile dye manufacturer developed a new yellow dye. The dye has a percent composition of 75.95% C, 17.72% N, and 6.33% H by mass with a molar mass of about 240 g/mol. Determine the molecular formula of the dye.

1. Copper(II) sulfide is oxidized by molecular oxygen to produce gaseous sulfur trioxide and solid copper(II) oxide. The gaseous product then reacts with liquid water to produce liquid hydrogen sulfate as the only product. Write the two equations which represent these reactions.

2. Write balanced chemical equations for the reactions used to prepare each of the following compounds from the given starting material(s). In some cases, additional reactants may be required.

(a) solid ammonium nitrate from gaseous molecular nitrogen via a two-step process (first reduce the nitrogen to ammonia, then neutralize the ammonia with an appropriate acid)

(b) gaseous hydrogen bromide from liquid molecular bromine via a one-step redox reaction

(c) gaseous H_2S from solid Zn and S via a two-step process (first a redox reaction between the starting materials, then reaction of the product with a strong acid)

3. Calcium cyclamate $Ca(C_6H_{11}NHSO_3)_2$ is an artificial sweetener used in many countries around the world but is banned in the United States. It can be purified industrially by converting it to the barium salt through reaction of the acid $C_6H_{11}NHSO_3H$ with barium carbonate, treatment with sulfuric acid (barium sulfate is very insoluble), and then neutralization with calcium hydroxide. Write the balanced equations for these reactions.

4. Complete and balance each of the following half-reactions (steps 2–5 in half-reaction method):

(a) $Sn^{4+}(aq) \longrightarrow Sn^{2+}(aq)$

(b) $Ag(NH_3)_2^+(aq) \longrightarrow Ag(s) + NH_3(aq)$

(c) $Hg_2Cl_2(s) \longrightarrow Hg(l) + Cl^-(aq)$

5. Complete and balance each of the following half-reactions (steps 2–5 in half-reaction method):

(a) $Cr^{2+}(aq) \longrightarrow Cr^{3+}(aq)$

(b) $Hg(l) + Br^-(aq) \longrightarrow HgBr_4^{2-}(aq)$

(c) $ZnS(s) \longrightarrow Zn(s) + S^{2-}(aq)$

(d) $Sn^{2+}(aq) + Cu^{2+}(aq) \longrightarrow Sn^{4+}(aq) + Cu^+(aq)$

(e) (h) $Fe^{3+}(aq) + I^-(aq) \longrightarrow Fe^{2+}(aq) + I_2(aq)$

6.5 Reaction Stoichiometry

6. Write the balanced equation, then outline the steps necessary to determine the information requested in each of the following:

(a) The number of moles and the mass of chlorine, Cl_2, required to react with 10.0 g of sodium metal, Na, to produce sodium chloride, NaCl.

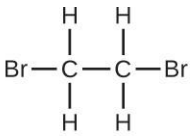
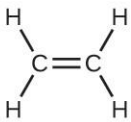
(b) The number of moles and the mass of oxygen formed by the decomposition of 1.252 g of mercury(II) oxide.

(c) The number of moles and the mass of sodium nitrate, $NaNO_3$, required to produce 128 g of oxygen. ($NaNO_2$ is the other product.)

(d) The number of moles and the mass of carbon dioxide formed by the combustion of 20.0 kg of carbon in an excess of oxygen.

(e) The number of moles and the mass of copper(II) carbonate needed to produce 1.500 kg of copper(II) oxide. (CO_2 is the other product.)

(f)

The number of moles and the mass of  formed by the reaction of 12.85 g of  with an excess of Br_2.

7. Determine the number of moles and the mass requested for each reaction in **Exercise 6.42**.

8. Write the balanced equation, then outline the steps necessary to determine the information requested in each of the following:

(a) The number of moles and the mass of Mg required to react with 5.00 g of HCl and produce $MgCl_2$ and H_2.

(b) The number of moles and the mass of oxygen formed by the decomposition of 1.252 g of silver(I) oxide.

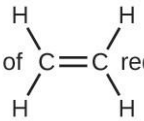
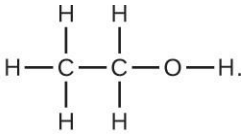
(c) The number of moles and the mass of magnesium carbonate, $MgCO_3$, required to produce 283 g of carbon dioxide. (MgO is the other product.)

(d) The number of moles and the mass of water formed by the combustion of 20.0 kg of acetylene, C_2H_2, in

an excess of oxygen.

(e) The number of moles and the mass of barium peroxide, BaO_2, needed to produce 2.500 kg of barium oxide, BaO (O_2 is the other product.)

(f)

The number of moles and the mass of  required to react with H_2O to produce 9.55 g of  .

9. I_2 is produced by the reaction of 0.4235 mol of $CuCl_2$ according to the following equation:

$$2CuCl_2 + 4KI \longrightarrow 2CuI + 4KCl + I_2.$$

(a) How many molecules of I_2 are produced?

(b) What mass of I_2 is produced?

10. What mass of silver oxide, Ag_2O, is required to produce 25.0 g of silver sulfadiazine, $AgC_{10}H_9N_4SO_2$, from the reaction of silver oxide and sulfadiazine?

$$2C_{10}H_{10}N_4SO_2 + Ag_2O \longrightarrow 2AgC_{10}H_9N_4SO_2 + H_2O$$

11. Carborundum is silicon carbide, SiC, a very hard material used as an abrasive on sandpaper and in other applications. It is prepared by the reaction of pure sand, SiO_2, with carbon at high temperature. Carbon monoxide, CO, is the other product of this reaction. Write the balanced equation for the reaction, and calculate how much SiO_2 is required to produce 3.00 kg of SiC.

12. Automotive air bags inflate when a sample of sodium azide, NaN_3, is very rapidly decomposed.

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

What mass of sodium azide is required to produce 2.6 ft^3 (73.6 L) of nitrogen gas with a density of 1.25 g/L?

13. Urea, $CO(NH_2)_2$, is manufactured on a large scale for use in producing urea-formaldehyde plastics and as a fertilizer. What is the maximum mass of urea that can be manufactured from the CO_2 produced by combustion of 1.00×10^3 kg of carbon followed by the reaction?

$$CO_2(g) + 2NH_3(g) \longrightarrow CO(NH_2)_2(s) + H_2O(l)$$

14. A compact car gets 37.5 miles per gallon on the highway. If gasoline contains 84.2% carbon by mass and has a density of 0.8205 g/mL, determine the mass of carbon dioxide produced during a 500-mile trip (3.785 liters per gallon).

6.6 Reaction Yields

15. The following quantities are placed in a container: 1.5×10^{24} atoms of hydrogen, 1.0 mol of sulfur, and 88.0 g of diatomic oxygen.

(a) What is the total mass in grams for the collection of all three elements?

(b) What is the total number of moles of atoms for the three elements?

(c) If the mixture of the three elements formed a compound with molecules that contain two hydrogen atoms,

one sulfur atom, and four oxygen atoms, which substance is consumed first?

(d) How many atoms of each remaining element would remain unreacted in the change described in (c)?

16. What is the limiting reactant in a reaction that produces sodium chloride from 8 g of sodium and 8 g of diatomic chlorine?

17. Which of the postulates of Dalton's atomic theory explains why we can calculate a theoretical yield for a chemical reaction?

18. A student isolated 25 g of a compound following a procedure that would theoretically yield 81 g. What was his percent yield?

19. A sample of 0.53 g of carbon dioxide was obtained by heating 1.31 g of calcium carbonate. What is the percent yield for this reaction?

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

20. Freon-12, CCl_2F_2, is prepared from CCl_4 by reaction with HF. The other product of this reaction is HCl. Outline the steps needed to determine the percent yield of a reaction that produces 12.5 g of CCl_2F_2 from 32.9 g of CCl_4. Freon-12 has been banned and is no longer used as a refrigerant because it catalyzes the decomposition of ozone and has a very long lifetime in the atmosphere. Determine the percent yield.

21. Citric acid, $C_6H_8O_7$, a component of jams, jellies, and fruity soft drinks, is prepared industrially via fermentation of sucrose by the mold *Aspergillus niger*. The equation representing this reaction is

$$C_{12}H_{22}O_{11} + H_2O + 3O_2 \longrightarrow 2C_6H_8O_7 + 4H_2O$$

What mass of citric acid is produced from exactly 1 metric ton (1.000×10^3 kg) of sucrose if the yield is 92.30%?

22. Toluene, $C_6H_5CH_3$, is oxidized by air under carefully controlled conditions to benzoic acid, $C_6H_5CO_2H$, which is used to prepare the food preservative sodium benzoate, $C_6H_5CO_2Na$. What is the percent yield of a reaction that converts 1.000 kg of toluene to 1.21 kg of benzoic acid?

$$2C_6H_5CH_3 + 3O_2 \longrightarrow 2C_6H_5CO_2H + 2H_2O$$

23. In a laboratory experiment, the reaction of 3.0 mol of H_2 with 2.0 mol of I_2 produced 1.0 mol of HI. Determine the theoretical yield in grams and the percent yield for this reaction.

24. Outline the steps needed to solve the following problem, then do the calculations. Ether, $(C_2H_5)_2O$, which was originally used as an anesthetic but has been replaced by safer and more effective medications, is prepared by the reaction of ethanol with sulfuric acid.

$2C_2H_5OH + H_2SO_4 \longrightarrow (C_2H_5)_2O + H_2SO_4{\cdot}H_2O$

What is the percent yield of ether if 1.17 L (d = 0.7134 g/mL) is isolated from the reaction of 1.500 L of C_2H_5OH (d = 0.7894 g/mL)?

25. Outline the steps needed to determine the limiting reactant when 30.0 g of propane, C_3H_8, is burned with 75.0 g of oxygen.

Determine the limiting reactant.

26. Outline the steps needed to determine the limiting reactant when 0.50 mol of Cr and 0.75 mol of H_3PO_4 react according to the following chemical equation.

$$2Cr + 2H_3PO_4 \longrightarrow 2CrPO_4 + 3H_2$$

Determine the limiting reactant.

27. What is the limiting reactant when 1.50 g of lithium and 1.50 g of nitrogen combine to form lithium nitride, a component of advanced batteries, according to the following unbalanced equation?

$$Li + N_2 \longrightarrow Li_3N$$

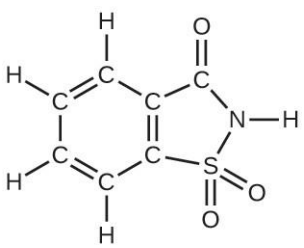
28. Uranium can be isolated from its ores by dissolving it as $UO_2(NO_3)_2$, then separating it as solid $UO_2(C_2O_4)\cdot 3H_2O$. Addition of 0.4031 g of sodium oxalate, $Na_2C_2O_4$, to a solution containing 1.481 g of uranyl nitrate, $UO_2(NO_3)_2$, yields 1.073 g of solid $UO_2(C_2O_4)\cdot 3H_2O$.

$Na_2C_2O_4 + UO_2(NO_3)_2 + 3H_2O \longrightarrow UO_2(C_2O_4)\cdot 3H_2O + 2NaNO_3$

Determine the limiting reactant and the percent yield of this reaction.

29. How many molecules of $C_2H_4Cl_2$ can be prepared from 15 C_2H_4 molecules and 8 Cl_2 molecules?

30. How many molecules of the sweetener saccharin can be prepared from 30 C atoms, 25 H atoms, 12 O atoms, 8 S atoms, and 14 N atoms?



31. The phosphorus pentoxide used to produce phosphoric acid for cola soft drinks is prepared by burning phosphorus in oxygen.

(a) What is the limiting reactant when 0.200 mol of P_4 and 0.200 mol of O_2 react according to

$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

(b) Calculate the percent yield if 10.0 g of P_4O_{10} is isolated from the reaction.

32. Would you agree to buy 1 trillion (1,000,000,000,000) gold atoms for $5? Explain why or why not. Find the current price of gold at http://money.cnn.com/data/commodities/ (1 troy ounce = 31.1 g)

6.7 Quantitative Chemical Analysis

33. A sample of gallium bromide, $GaBr_3$, weighing 0.165 g was dissolved in water and treated with silver nitrate, $AgNO_3$, resulting in the precipitation of 0.299 g AgBr. Use these data to compute the %Ga (by mass) $GaBr_3$.

34. The principal component of mothballs is naphthalene, a compound with a molecular mass of about 130 amu, containing only carbon and hydrogen. A 3.000-mg sample of naphthalene burns to give 10.3 mg of CO_2. Determine its empirical and molecular formulas.

35. A 0.025-g sample of a compound composed of boron and hydrogen, with a molecular mass of ~28 amu, burns spontaneously when exposed to air, producing 0.063 g of B_2O_3. What are the empirical and molecular formulas of the compound?

36. Sodium bicarbonate (baking soda), $NaHCO_3$, can be purified by dissolving it in hot water (60 °C), filtering to remove insoluble impurities, cooling to 0 °C to precipitate solid $NaHCO_3$, and then filtering to remove the solid, leaving soluble impurities in solution. Any $NaHCO_3$ that remains in solution is not recovered. The solubility of $NaHCO_3$ in hot water of 60 °C is 164 g/L. Its solubility in cold water of 0 °C is 69 g/L. What is the percent yield of $NaHCO_3$ when it is purified by this method?

37. The reaction of WCl_6 with Al at ~400 °C gives black crystals of a compound containing only tungsten and chlorine. A sample of this compound, when reduced with hydrogen, gives 0.2232 g of tungsten metal and hydrogen chloride, which is absorbed in water. Titration of the hydrochloric acid thus produced requires 46.2 mL of 0.1051 *M* NaOH to reach the end point. What is the empirical formula of the black tungsten chloride?