1.0 INTRODUCTION

Chemistry - A Molecular Science (CAMS), the first half of this two-volume sequence, stressed bonding, structure, and reactivity. The material was qualitative and stressed several types of reactions and the factors that affected their relative extents of reaction. However, as the title of this text suggests, chemistry is also a quantitative science. Chemists must not only predict the products of a reaction, they must also predict the amount of product that can be expected, and the amount of waste that must be removed. They also need to know how much energy is required or how much heat is generated by a reaction. They must also understand how the reaction occurs so that they can optimize the reaction conditions. These are the types of problems addressed in this text.

We begin our study of the quantitative aspects of chemistry with **stoichiometry**, the science that deals with the quantitative relationships between the elements in a compound (substance stoichiometry) and between the substances in a chemical reaction (reaction stoichiometry). It is the topic of this first chapter because a thorough knowledge of stoichiometry is vital to an understanding of the material presented in this course. Understanding how quantitative data and results are presented is also important, so you should review *Appendix A, Reporting Quantitative Measurements and Results*, for a treatment of precision, significant figures, and rounding errors.

**THE OBJECTIVES OF THIS CHAPTER ARE TO SHOW YOU HOW TO:**

- determine the formula of a substance from its composition;
- balance chemical equations by inspection;
- use chemical equations to determine the relative amounts of reactants and products involved in a reaction;
- determine a limiting reactant; and
- determine the composition of a reaction mixture after the reaction is complete.
1.1 THE MOLE

Chemists use chemical equations to design possible routes to desired molecules and to discuss chemical processes. However, the individual molecules represented in the equations are far too small to be seen, so chemists must use a very large number of molecules in their reaction in order that the reactants and products can be observed. Indeed, the number of molecules required to make a visible sample is staggering. Consider that 1 μL of water - about 1/50th of a drop - contains about 100,000,000,000,000,000 or \(10^{17}\) molecules and a typical reaction in the laboratory involves thousands of times that number. Such large numbers are cumbersome, so scientists use a more convenient unit when discussing numbers of molecules. This unit is called a mole (mol):

\[
1 \text{ mol} = 6.0221\times10^{23} \text{ items}
\]

6.0221\times10^{23} = \(N_A\) is Avogadro’s number. A mole is used to indicate a number of atoms just as a dozen is used to indicate a number of eggs. Converting from moles to atoms is done the same as converting dozens to items. 1.5 doz = (1.5 doz)(12 items·doz\(^{-1}\)) = 18 items and 1.5 mol = (1.5 mol)( 6.0\times10^{23} \text{ atoms·mol}^{-1}) = 9.0\times10^{23} \text{ atoms}. The mole is used simply because it is much easier to discuss the number of atoms in moles than it is as individual items - 0.10 mol H\(_2\)O is a much more convenient expression than 6.0\times10^{22} \text{ H}_2\text{O} molecules.

Chemists need to be able to readily prepare mixtures of reactants that have the correct atom or molecule ratios to react, but they certainly cannot count such large numbers. Instead, they use other more easily determined properties that are related to the numbers of atoms and/or molecules. The first such method we examine is mass. Mass can be used to ‘count’ atoms and molecules because a mole is the number of atoms present in one gram atomic weight* of any atom or in one gram molecular weight of any molecule. Thus, the mass of a mole of any substance, which is known as its molar mass (\(M_m\)), equals its atomic or molecular weight expressed in grams. For example, the atomic weight of Mg is 24.3, so its molar mass is 24.3 g·mol\(^{-1}\), and the molecular weight of CO\(_2\) is 44.0, so its molar mass is 44.0 g·mol\(^{-1}\). Thus, molar mass allows us to quickly convert a mass into a number of moles or a number of moles into a mass. Chemists use this fact to quickly ‘count’ the number of moles of substance by simply weighing it.

Mass ↔ mole conversions are most easily done with the factor-label method. It uses the units of the given quantity and those of the conversion factors to assure the proper operations are performed. To use this method, arrange the factors so that the denominator

\* A gram atomic weight is a mass of atoms equal to the atomic weight of the atom expressed in grams.
of each factor cancels the numerator of the previous quantity until the units of the answer are obtained. This is shown explicitly in the following examples, where the units that cancel have lines drawn through them.

### Example 1.1

Determine the molar masses to the nearest whole number.

**N₂F₄**

One mole of N₂F₄ contains 2 mol N and 4 mol F. The atomic masses of N and F are 14 and 19, respectively, so the molar mass of N₂F₄ is

\[
\frac{2 \text{ mol N}}{1 \text{ mol N}_2\text{F}_4} \times \frac{14 \text{ g N}}{1 \text{ mol N}} + \frac{4 \text{ mol F}}{1 \text{ mol N}_2\text{F}_4} \times \frac{19 \text{ g F}}{1 \text{ mol F}} = \frac{104 \text{ g N}_2\text{F}_4}{1 \text{ mol N}_2\text{F}_4} = 104 \text{ g mol}^{-1}
\]

We included the ‘per mol substance’ in the above, but it will be implied rather than written explicitly in future molar mass determinations.

**Ca₃(PO₄)₂**

One mol Ca₃(PO₄)₂ contains 3 mol Ca, 2 mol P, and 8 mol O, so the mass of one mole is

\[
\frac{3 \text{ mol Ca}}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times \frac{40 \text{ g Ca}}{1 \text{ mol Ca}} + \frac{2 \text{ mol P}}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times \frac{31 \text{ g P}}{1 \text{ mol P}} + \frac{8 \text{ mol O}}{1 \text{ mol Ca}_3(\text{PO}_4)_2} \times \frac{16 \text{ g O}}{1 \text{ mol O}} = 310 \text{ g mol}^{-1}
\]

The ‘per mol Ca₃(PO₄)₂’ is not included, but it is implied in the molar mass.

### Example 1.2

**a)** What is the mass of 3.24 mol N₂O₅?

First, use the molar masses of N and O to determine the molar mass of N₂O₅.

\[
\frac{2 \text{ mol N}}{1 \text{ mol N}_2\text{O}_5} \times \frac{14.0 \text{ g N}}{1 \text{ mol N}} + \frac{5 \text{ mol O}}{1 \text{ mol O}_2} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = 108.0 \text{ g mol}^{-1}
\]

Next, use the molar mass to convert the given moles into mass.

\[
3.24 \text{ mol N}_2\text{O}_5 \times \frac{108.0 \text{ g N}_2\text{O}_5}{1 \text{ mol N}_2\text{O}_5} = 350. \text{ g N}_2\text{O}_5
\]

**b)** How many moles of N₂O₅ are present in a 12.7-g sample of N₂O₅?

Use the molar mass determined in Part A to convert from mass to moles.

\[
\frac{12.7 \text{ g N}_2\text{O}_5}{108.0 \text{ g N}_2\text{O}_5} = 0.118 \text{ mol N}_2\text{O}_5
\]

### PRACTICE EXAMPLE 1.1

Determine the masses of the following.

**a)** 2.88 mol PF₃

**mass**

\[
M_m = \text{_________ g mol}^{-1}
\]

**b)** 0.0448 mol C₃H₈O

**mass**

\[
M_m = \text{_________ g mol}^{-1}
\]

Determine the number of moles of compound in the following.

**c)** 18.6 g K₂SO₄

**moles**

\[
M_m = \text{_________ g mol}^{-1}
\]

**d)** 0.2668 g H₃PO₄

**moles**

\[
M_m = \text{_________ g mol}^{-1}
\]
Example 1.3

How many Al atoms are present in a piece of aluminum foil that has a mass 0.065 g?

The number of atoms is given by the number of moles of Al (M_m = 27 g·mol⁻¹).

\[
\frac{0.065 \text{ g Al}}{27 \text{ g Al}} \times \frac{1 \text{ mol Al}}{27 \text{ g Al}} = 2.4 \times 10^{-3} \text{ mol Al}
\]

which is a perfectly good answer to the question. However, the number of moles of Al can be converted to the number atoms with the use of Avogadro’s number.

\[
2.4 \times 10^{-3} \text{ mol Al} \times \frac{6.02 \times 10^{23} \text{ Al atoms}}{1 \text{ mol Al}} = 1.4 \times 10^{21} \text{ Al atoms}
\]

Both steps can be combined into one operation as follows:

\[
\frac{0.065 \text{ g Al}}{27 \text{ g Al}} \times \frac{1 \text{ mol Al}}{27 \text{ g Al}} \times \frac{6.02 \times 10^{23} \text{ Al atoms}}{1 \text{ mol Al}} = 1.4 \times 10^{21} \text{ Al atoms}
\]

The number of moles of molecules in a gas can also be determined with the ideal gas law.

\[
P \cdot V = n \cdot R \cdot T \quad \text{Eq. 1.1}
\]

P is the pressure of the gas in atmospheres, V is its volume in liters, n is the number of moles of gas, R = 0.08206 L·atm·K⁻¹·mol⁻¹ is the ideal gas law constant, and T is the temperature on the Kelvin scale (K = °C + 273.15).

Example 1.4

How many moles of H₂ are in a 3.06 L container at 22°C if its pressure is 742 torr?

Convert the Celsius temperature to the Kelvin scale: T = 22 + 273 = 293 K

Use the equality 760 torr = 1 atm to convert the pressure to atmospheres.

\[
P = \frac{742 \text{ torr}}{760 \text{ torr}} = 0.976 \text{ atm}
\]

Solve the ideal gas law for n and substitute the known quantities.

\[
n = \frac{PV}{RT} = \frac{(0.976 \text{ atm})(3.06 \text{ L})}{(0.08206 \text{ L·atm·K}^{-1}·\text{mol}^{-1})(293 \text{ K})} = 0.124 \text{ mol}
\]

PRACTICE EXAMPLE 1.2

How many molecules are present in each sample?

a) 2.66 mmol CO₂

\[\underline{\quad \text{molecules of CO₂}}\]

b) 12.0 μg of N₂O₅

\[\underline{\quad \text{molecules of N₂O₅}}\]

PRACTICE EXAMPLE 1.3

What is the mass of CO₂ in a 500.0 mL flask at 75°C if its pressure is 1089 torr?

Pressure in atmospheres

\[P = \underline{\quad \text{atm}}\]

Temperature on the Kelvin scale

\[T = \underline{\quad \text{K}}\]

Moles of CO₂

\[n = \underline{\quad \text{mol}}\]

Mass of CO₂

\[m = \underline{\quad \text{g}}\]
1.2 DETERMINING CHEMICAL FORMULAS

The **elemental composition** of a substance is typically given as the mass percents of its component elements. The **mass percent** of an element in a compound is the fraction of the total mass of the compound due to the element expressed as a percent (part of a hundred).

**Example 1.5**

A 3.17-g sample of an oxide of lead was found to contain 2.94 g of lead. What is the elemental composition of the oxide expressed as mass percents?

The mass of sample and the mass of lead in the sample are given, but the mass of oxygen must be determined by difference.

\[
\text{Mass of O} = \text{mass of sample} - \text{mass of Pb} = 3.17 - 2.94 = 0.23 \text{ g O}
\]

The mass percent of each element is determined as the mass of the element divided by the mass of the sample times 100%.

\[
\frac{2.94 \text{ g Pb}}{3.17 \text{ g oxide}} \times 100\% = 92.7\% \text{ Pb} \quad \text{and} \quad \frac{0.23 \text{ g O}}{3.17 \text{ g oxide}} \times 100\% = 7.3\% \text{ O}
\]

Alternatively, we could have found the mass percent of Pb with the given data and then used the fact that the sum of the mass percents of all elements in the compound must sum to 100%. Thus, the mass percent of O could be found as follows:

\[
\% \text{O} = 100.0\% \text{ total} - 92.7\% \text{ Pb} = 7.3\% \text{ O}
\]

In Example 1.5, the elemental composition of a compound was determined from experimental data. However, elemental compositions can also be determined from the chemical formula and molar masses. Consider the case of Fe₂O₃, the material responsible for the orange color of clay. One mole of Fe₂O₃ has a mass of

\[
2 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} + 3 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 111.70 \text{ g Fe} + 48.00 \text{ g O} = 159.70 \text{ g Fe₂O₃}
\]

A mole of Fe₂O₃ has a mass of 159.70 g and contains 111.70 g Fe and 48.00 g of O. To determine the elemental composition of a substance as mass fraction, divide each of the elemental masses by the molar mass of the substance,

\[
\text{mass fraction Fe} = \frac{111.70 \text{ g Fe}}{159.70 \text{ g Fe₂O₃}} = 0.6994; \quad \text{mass fraction O} = \frac{48.00 \text{ g O}}{159.70 \text{ g Fe₂O₃}} = 0.3006
\]

Multiplication of the mass fraction by 100 converts it into a **mass percent**. Thus, any pure sample of Fe₂O₃ is 69.94% iron and 30.06% oxygen by mass. Mass percents are also conversion factors that can be used to convert between a mass of a compound and the
masses of its elements because the units of the mass percent of an element can be expressed as (g of element/100 g compound).

### Example 1.6

a) Ammonium nitrate is a good source of nitrogen that is used in the fertilizer industry. What percent of the mass of NH₄NO₃ is due to nitrogen?

Determine the molar mass of NH₄NO₃

\[
M_m = 2 \text{ mol N} \times \frac{14.0 \text{ g N}}{1 \text{ mol N}} + 4 \text{ mol H} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} + 3 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = 80.0 \text{ g NH₄NO₃}
\]

80.0 g of ammonium nitrate contains 28.0 g of nitrogen, so the percent nitrogen is

\[
% \text{ N} = \frac{28.0 \text{ g N}}{80.0 \text{ g NH₄NO₃}} \times 100\% = 35.0\%
\]

b) How many pounds of N are in 25.0 lb of NH₄NO₃?

Mass percent is a ratio of masses, so the choice of mass units in the factor is arbitrary as long as they are the same in the numerator and the denominator. Thus, we can use the above mass percent to convert pounds even though it was determined from grams.

\[
25.0 \text{ pounds NH₄NO₃} \times \frac{35.0 \text{ pounds N}}{100 \text{ pounds NH₄NO₃}} = 8.75 \text{ pounds N}
\]

Alternatively, we could use the ratio used to get the percents as our conversion factor

\[
25.0 \text{ pounds NH₄NO₃} \times \frac{28 \text{ pounds N}}{80.0 \text{ pounds NH₄NO₃}} = 8.75 \text{ pounds N}
\]

The formula Fe₂O₃ shows two iron atoms for every three oxygen atoms; it does not indicate that a molecule of Fe₂O₃ has two iron atoms and three oxygen atoms. Indeed, Fe₂O₃, which is the mineral hematite, exists as an extended solid with no discreet Fe₂O₃ units. Formulas that show only the smallest whole number ratio of atoms present in a compound are called simplest formulas or empirical formulas. The simplest or empirical formula of a substance can be determined from its mass composition in two steps:

1. Determine the number of moles of each element present in a fixed mass of the compound. This information can be given as experimental masses as in Example 1.5, or as percent compositions as in Example 1.6. If percent composition is given, the simplest procedure is to assume a fixed mass of 100 g, so the percents equal the masses of the elements present.

### PRACTICE EXAMPLE 1.5

What are the mass percents of the elements in Ca(ClO₃)₂?

Molar mass of Ca(ClO₃)₂ = _____________ g·mol⁻¹

mass % Ca = x100% = _________%

mass % Cl = x100% = _________%

mass % O = x100% = _________%
2. Find the simplest whole number ratio of the moles present by dividing each result of Step 1 by the smallest result. If any of the new numbers are not integers, multiply all of the numbers by the number that makes them all integers. The resulting integers are the subscripts of the simplest formula. Table 1.1 gives the multipliers to use for some common decimals.

Example 1.7

A 7.50-g sample of iron is heated in oxygen to form an iron oxide. If 10.36 g of the oxide is formed, what is its simplest formula?

In this example, the elemental masses are given. The fixed mass of the compound is 10.36 g, and it contains 7.50 g Fe. The mass of oxygen that it contains is determined by difference: mass O = 10.36 g Fe\textsubscript{2}O\textsubscript{y} – 7.50 g Fe = 2.86 g O.

Step 1. Convert the two elemental masses to moles.

\[
7.50 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.134 \text{ mol Fe} \quad \text{and} \quad 2.86 \text{ g O} \times \frac{1 \text{ mol O}}{16.0 \text{ g O}} = 0.179 \text{ mol O}
\]

Step 2. Determine the ratio of the elemental moles obtained in Step 1.

\[
\frac{0.179 \text{ mol O}}{0.134 \text{ mol Fe}} = \frac{1.33 \text{ mol O}}{1 \text{ mol Fe}}
\]

Multiply the numerator and denominator by 3 to obtain a ratio of integers.

\[
\frac{(3)(1.33 \text{ mol O})}{(3)(1 \text{ mol Fe})} = \frac{4 \text{ mol O}}{3 \text{ mol Fe}}
\]

There are 4 mol O for every 3 mol Fe, so the simplest formula is Fe\textsubscript{3}O\textsubscript{4}, which is the mineral known as magnetite.

Example 1.8

KClO\textsubscript{x} decomposes into KCl and O\textsubscript{2}. What is the value of x in a compound that produces 468.5 mg KCl and 230.0 mL of O\textsubscript{2} gas at 756.2 torr and 23 °C?

X is the number of moles of O per mole of KCl, so we find the number of moles of each. Use the ideal gas law to determine moles of O\textsubscript{2}.

\[
n = \frac{PV}{RT} = \frac{(756/760 \text{ atm})(0.230 \text{ L})}{(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(296 \text{ K})} = 9.42 \times 10^{-3} \text{ mol} = 9.42 \text{ mmol}
\]

There are 2 mol O/mol O\textsubscript{2}, so the sample contains 2(9.42) = 18.84 mmol O. Divide the given mass of KCl by its molar mass (74.6 g/mol) to obtain moles of KCl.

Table 1.1 Some common decimals and their multipliers

<table>
<thead>
<tr>
<th>decimal</th>
<th>multiplier</th>
<th>decimal</th>
<th>multiplier</th>
</tr>
</thead>
<tbody>
<tr>
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<td>0.5</td>
<td>2</td>
</tr>
<tr>
<td>0.167</td>
<td>6</td>
<td>0.625</td>
<td>8</td>
</tr>
<tr>
<td>0.200</td>
<td>5</td>
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<td>3</td>
</tr>
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<td>0.250</td>
<td>4</td>
<td>0.75</td>
<td>4</td>
</tr>
<tr>
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<td>3</td>
<td>0.833</td>
<td>6</td>
</tr>
<tr>
<td>0.375</td>
<td>8</td>
<td>0.875</td>
<td>8</td>
</tr>
</tbody>
</table>

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Chapter 1  Stoichiometry

8

\[ 0.4685 \text{ g KCl} \times \frac{1 \text{ mol KCl}}{74.6 \text{ g KCl}} = 6.28 \times 10^{-3} \text{ mol} = 6.28 \text{ mmol} \]

Obtain \( x \) as the ratio of moles O/moles KCl

\[ x = \frac{18.84 \text{ mmol}}{6.28 \text{ mmol}} = 3.00 \]

The compound is KClO₃, potassium chlorate.

The empirical formulas of benzene and acetylene are both CH. Indeed, there are hundreds of compounds with that empirical formula. Yet, there are no molecules that are composed of a single carbon atom and a single hydrogen atom. Formulas that represent the actual numbers of atoms in a molecule are called **molecular formulas**. A molecular formula always contains an integral number of simplest or empirical formulas: molecular formula = (simplest formula)ₙ. The molecular formula of benzene is C₆H₆, so a benzene molecule contains six empirical units, (CH)₆. The molar mass of the compound must also be an integral number of simplest formula molar masses: \( M_{m(\text{compound})} = nM_{m(\text{empirical formula})} \), where \( n \) is an integer. Thus, the value of \( n \) and the molecular formula can be determined from the empirical formula if the molar mass of the compound is known.

\[ n = \frac{\text{molar mass of molecular formula}}{\text{molar mass of empirical formula}} \quad \text{Eq. 1.2} \]

For example, the molar mass of the CH unit is 13 g·mol⁻¹ and the molar mass of C₆H₆ is 78 g·mol⁻¹, so we would determine \( n \) as follows:

\[ n = \frac{78 \text{ g mol}^{-1}}{13 \text{ g mol}^{-1}} = 6 \]

**Example 1.9**

a) The amino acid lysine is 49.296% C, 9.653% H, 19.162% N, and 21.889% O. What is the empirical or simplest formula of lysine?

Step 1. Assume a 100-g sample of lysine, so the mass of each element is equal to its percent then determine the number of moles of each element.

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass (g)</th>
<th>Molar Mass (g/mol)</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>49.296</td>
<td>12.011</td>
<td>4.104</td>
</tr>
<tr>
<td>H</td>
<td>9.653</td>
<td>1.008</td>
<td>9.576</td>
</tr>
<tr>
<td>N</td>
<td>19.162</td>
<td>14.007</td>
<td>1.368</td>
</tr>
<tr>
<td>O</td>
<td>21.889</td>
<td>15.999</td>
<td>1.368</td>
</tr>
</tbody>
</table>

b) What is the value of \( x \) (mmols H₂O/mmol CoSO₄) in the formula of the hydrate?

\[ M_{m(\text{H}_2\text{O})} = \text{mg mmol}^{-1} \]

\[ M_{m(\text{CoSO}_4)} = \text{mg mmol}^{-1} \]

The hydrate contains

\[ x = \frac{\text{mmol H}_2\text{O}}{\text{mmol CoSO}_4} = \text{mmol CoSO}_4 \]

The formula of the hydrate is __________
Example 1.10

a) What is the empirical formula of ascorbic acid (vitamin C) if combustion of a 0.579-g sample of ascorbic acid produced 0.868 g CO₂ and 0.237 g H₂O.

Ascorbic acid contains only C, H, and O atoms. Combustion converts all of the carbon into CO₂ and all of the hydrogen into H₂O, so the number of moles of carbon and hydrogen in the sample is determined as follows:

\[
\frac{0.868 \text{ g CO}_2}{44.01 \text{ g CO}_2} \times 1 \text{ mol C} = 0.0197 \text{ mol C}
\]
\[
\frac{0.237 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times 2 \text{ mol H} = 0.0263 \text{ mol H}
\]

Oxygen is added in the combustion reaction, so the amount in the sample cannot be determined from the amounts of CO₂ and H₂O. Instead, we must determine the mass of oxygen by difference. Thus, we convert the moles of carbon and hydrogen into grams.

\[
0.0197 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.237 \text{ g C}
\]
\[
0.0263 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0265 \text{ g H}
\]

Then we use the total mass of the sample and the masses of the elements to get the mass of oxygen by difference.

\[
\text{mass O} = 0.579 \text{ g ascorbic acid} - 0.237 \text{ g C} - 0.0265 \text{ g H} = 0.316 \text{ g O}
\]

b) The molar mass of lysine is 146.18 g mol⁻¹, what is its molecular formula?

The molar mass of an empirical unit of lysine is: 3(12.011) g mol⁻¹ C + 7(1.008) g mol⁻¹ H + 14.007 g mol⁻¹ N + 15.999 g mol⁻¹ O = 73.090 g mol⁻¹, so n is determined to be

\[
n = \frac{M_{\text{m}}(\text{lysine})}{M_{\text{m}}(\text{C}_3\text{H}_7\text{NO})} = \frac{146.18 \text{ g mol}^{-1}}{73.090 \text{ g mol}^{-1}} = 2.000
\]

The molecular formula of lysine is (C₃H₇NO)₂. However, the molecular formula is written as = C₆H₁₄N₂O₂ because (C₃H₇NO)₂ incorrectly implies that lysine is composed of two identical C₃H₇NO units.*

\[
\text{H₂N} \begin{array}{c} \text{C} \text{C} \text{OH} \\ \text{H} \text{(CH₂)₄} \text{NH₂} \end{array}
\]

Lysine is an amino acid with a molecular formula of C₆H₁₄N₂O₂.

---

* We convert to masses only because we need to know how much oxygen the sample contains. If the sample did not contain oxygen, we would simply determine the mol H/mol C ratio from the moles determined above to get the empirical formula.
Determine the number of moles of oxygen in the sample

\[ 0.316 \text{ g } \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0197 \text{ mol O} \]

The number of moles of C and O are the same, so their subscripts are the same. To find the ratio of C or O to H, divide the number of moles of H by the number of moles of C.

\[ \frac{0.0263 \text{ mol H}}{0.0197 \text{ mol C}} = \frac{1.33 \text{ mol H}}{1 \text{ mol C}} \]

Multiplication of the numerator and denominator by 3 to eliminate the decimal yields a 4:3 ratio, and the empirical formula is \( \text{C}_3\text{H}_4\text{O}_3 \).

**b)** What is its molecular formula if its molar mass is 176 g·mol\(^{-1}\)?

The molar mass of the empirical unit is \( 3(12.01) + 4(1.01) + 3(16.00) = 88.1 \text{ g·mol}^{-1} \). Division of the molar mass of the compound by the molar mass of the empirical unit yields the number of empirical units in the molecular formula. \( \frac{176}{88} = 2 \), so the molecular formula is \( \text{C}_6\text{H}_8\text{O}_6 \). Again, the formula would not be written as \((\text{C}_3\text{H}_4\text{O}_3)^2\) because there are not two identical and identifiable \( \text{C}_3\text{H}_4\text{O}_3 \) units in the molecule.*

### 1.3 SUBSTANCE OR COMPOSITION STOICHIOMETRY

Stoichiometry problems involve the conversion of an amount of one substance (the given substance) into a comparable amount of another substance (the substance that is sought). This is done by converting the given amount to moles and then multiplying by the mole ratio that relates the sought and given substances.

\[
\text{moles given} \times \frac{\text{moles sought}}{\text{moles given}} = \text{moles sought}
\]

Eq. 1.3

The mole ratio is called the **stoichiometric factor, link, or ratio**. In substance or composition stoichiometry, the ratio is obtained from the subscripts in the chemical formula of the compound. The stoichiometric factors that can be obtained from the formula \( \text{Fe}_2\text{O}_3 \) are

\[
\begin{align*}
2 \text{ mol Fe} & \quad 3 \text{ mol O} & \quad 2 \text{ mol Fe} \\
1 \text{ mol Fe}_2\text{O}_3 & \quad 1 \text{ mol Fe}_2\text{O}_3 & \quad 3 \text{ mol O}
\end{align*}
\]

Note that the reciprocals of the above are also stoichiometric factors. As demonstrated in Example 1.11, the amount of one element that is combined with a known amount of another can be determined by using Equation 1.3 and molar masses.

---

*Ascorbic acid (vitamin C) has a molecular formula of \( \text{C}_6\text{H}_8\text{O}_6 \).

**PRACTICE EXAMPLE 1.7**

a) What is the simplest formula of a hydrocarbon if a 0.250-g sample produces 0.784 g of CO\(_2\) and 0.321 g of H\(_2\)O upon complete combustion?

The sample contains

\[
\begin{align*}
\text{mol H} & \quad \text{mol C} \\
\text{H}_2\text{O} & \quad \text{CO}_2
\end{align*}
\]

The simplest whole number mole ratio is

\[ \text{mol H/mol C} = \]

The empirical formula is __________

b) What is the molecular formula of the compound if its molar mass is 84.16 g/mol?

The number of empirical units in molecular formula

\[ n = \]

The molecular formula is ____________
Example 1.11

What is the mass of chlorine in a sample of CCl₄ that contains 4.72 g of carbon?

Convert the given mass of carbon into moles, then use stoichiometric ratio to convert moles of carbon to moles of chlorine. Finally, use the molar mass of chlorine to calculate the mass of Cl. The factor-label method can be used to establish the order of each operation: start with the given quantity and string the factors so that the denominator of each has the same unit as the previous numerator.

\[
4.72 \text{ g C} \times \frac{1 \text{ mol C}}{12.0 \text{ g C}} \times \frac{4 \text{ mol Cl}}{1 \text{ mol C}} \times \frac{35.5 \text{ g Cl}}{1 \text{ mol Cl}} = 55.9 \text{ g Cl}
\]

Example 1.11 is a typical stoichiometry problem. First, the mass of the given substance is converted into moles by dividing by its molar mass. Next, the moles of the given substance are multiplied by the stoichiometric factor to obtain the moles of the desired substance. In a composition stoichiometry problem, the stoichiometric factor is the ratio of the subscripts in the formula. Finally, the number of moles of the desired substance is converted to mass by multiplying by its molar mass. The process for determining the mass of reactant or product in a reaction is identical except that the stoichiometric ratio is obtained from a balanced equation, the topic of Section 1.4.

1.4 BALANCING CHEMICAL EQUATIONS

A chemical equation reads like a sentence, where the formulas of the reactants and the products are the words that are read from left to right. Thus, the reactants (substances present before reaction) are on the left while the products (substances present after reaction) appear on the right. A chemical equation expresses the relative amounts and the identities of the substances involved in chemical and physical changes.

Neither the number nor the identity of the atoms involved in a chemical reaction changes. Consequently, chemical equations are balanced to assure that the number of atoms of each kind is the same on both sides. A procedure that can be used to balance many chemical equations is demonstrated by balancing the following chemical equation:

\[
\text{Ca}_3\text{N}_2 + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{NH}_3
\]

**Step 1. Identify a starting substance.**

Start with the substance on either side of the equation that has the greatest number of atoms or the largest subscripts. In this case, both Ca₃N₂ and Ca(OH)₂ contain five atoms, but we start with Ca₂N₃ because its subscripts are larger.

PRACTICE EXAMPLE 1.8

What mass of barium is in a sample of Ba₃(PO₄)₂ that contains 4.00 g of phosphorus?

\[
\text{mass} = \underline{\text{________}} \text{ g Ba}
\]
Step 2. Identify a coefficient for the starting substance.
A coefficient of 1 (one) is used as the starting coefficient unless there is an obvious reason to choose a different number.

\[ \text{Step 3.} \quad \text{Determine which atoms have been fixed in Step 2. Then balance those atoms by placing coefficients on the other side of the equation.} \]

A coefficient of 1 for \( \text{Ca}_3\text{N}_2 \) sets the number of calcium and nitrogen atoms on the reactant (left) side at three and two, respectively, so the coefficients of \( \text{Ca(OH)}_2 \) and \( \text{NH}_3 \) must be three and two, respectively.

\[ \text{Step 4.} \quad \text{Balance those atoms that have been fixed in Step 3, by placing coefficients on the side of the equation opposite to the side worked on in Step 3.} \]

The number of hydrogen atoms and oxygen atoms were both fixed on the product side in Step 3. The 12 hydrogen atoms and 6 oxygen atoms are balanced by fixing the coefficient of water at 6.

The equation is now balanced because there are 3Ca, 2N, 12H, and 6O on each side. Coefficients of one are not usually included, so the reaction would be written as

\[ \text{Ca}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Ca(OH)}_2 + 2\text{NH}_3 \]

In a more complicated reaction, the above process, moving back and forth between the left and right sides of the reaction, is continued until all of the atoms are balanced.

Example 1.12

Balance the following chemical equation, which is the reaction for the commercial production of phosphorus:

\[ \text{Step 1.} \quad \text{We choose Mg}_3(\text{PO}_4)_2 \text{ as the starting substance because it contains the greatest number of atoms.} \]

Step 2. Note that a coefficient of one for \( \text{Mg}_3(\text{PO}_4)_2 \) produces only two \( P \) atoms, while a minimum of four are needed to balance the right side. Consequently, we use two as our starting coefficient.

\[ \text{Step 3.} \quad \text{Our starting coefficient has fixed the number of magnesium atoms at six and the number of phosphorus atoms at four. However, it did not fix the number of oxygen atoms} \]
because we do not yet know the coefficient of SiO₂. Balancing the magnesium and phosphorus atoms on the product side of the reaction, we obtain

\[ \underline{\text{C}} + \underline{\text{SiO}_2} + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + \underline{\text{CO}} \]

Step 4. The coefficient of MgSiO₃ fixes the number of silicon atoms at six. The oxygen atoms are not fixed because we do not yet know the coefficient of CO. Balancing the silicon atoms on the reactant side, we obtain

\[ \underline{\text{C}} + 6\text{SiO}_2 + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + \underline{\text{CO}} \]

Step 5. The number of oxygen atoms has now been fixed at 28 on the reactant side (16 in 2Mg₃(PO₄)₂ and 12 in 6SiO₂). There are already 18 oxygen atoms on the product side (6MgSiO₃), so only ten must be balanced with CO.

\[ \underline{\text{C}} + 6\text{SiO}_2 + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + 10\text{CO} \]

Step 6. Balance the carbon atoms to obtain the final balanced equation.

\[ 10\text{C} + 6\text{SiO}_2 + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + 10\text{CO} \]

The procedure presented above can be used to balance most chemical reactions. However, there are some redox reactions that cannot be balanced by inspection. The methods used to balance redox equations are presented in Appendix F.

### 1.5 REACTION STOICHIOMETRY

In reaction stoichiometry, the amount of one substance that reacts with, is produced by, or is required to produce a given amount of another substance is determined. The problems are done in a manner that is identical to the method shown for composition stoichiometry problems (Example 1.11) except for the nature of the stoichiometric ratio. In reaction stoichiometry, the stoichiometric ratios are derived from the coefficients in the balanced equation. For example, consider the reaction that was balanced in Example 1.12.

\[ 10\text{C} + 6\text{SiO}_2 + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + 10\text{CO} \]

Some of the stoichiometric ratios that can be derived from this reaction are

<table>
<thead>
<tr>
<th>Substance</th>
<th>Coefficient</th>
<th>Stoichiometric Ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>10 mol C</td>
<td>6 mol SiO₂</td>
<td>(\frac{10}{6}) or 1.67 mol</td>
</tr>
<tr>
<td>6 mol SiO₂</td>
<td>2 mol (\text{Mg}_3(\text{PO}_4)_2)</td>
<td>(\frac{6}{2}) or 3</td>
</tr>
<tr>
<td>2 mol (\text{Mg}_3(\text{PO}_4)_2)</td>
<td>1 mol (\text{P}_4)</td>
<td>(\frac{2}{1}) or 2</td>
</tr>
<tr>
<td>1 mol (\text{P}_4)</td>
<td>10 mol CO</td>
<td>(\frac{1}{10}) or 0.1</td>
</tr>
</tbody>
</table>

The stoichiometric ratio converts the number of moles of a given substance (denominator) into the equivalent number of moles of a desired substance (numerator).

The number of moles of substance can be given directly, but it is more often given by its mass if it is a solid, its pressure, volume and temperature if it is a gas, or by its volume and molarity if it is a solute. We consider only the first two types of calculations in this chapter and postpone examples using molarity until Chapter 2. However, once the moles...
of the given substance have been determined, they are multiplied by the appropriate stoichiometric ratio regardless of how they were determined. Example 1.13 illustrates the procedure.

Example 1.13

All of the following problems are based on the reaction balanced in Example 1.12:

\[
10C + 6SiO_2 + 2Mg_3(PO_4)_2 \rightarrow P_4 + 6MgSiO_3 + 10CO
\]

a) What minimum mass of SiO₂ (\(M_m = 60.1 \text{ g·mol}^{-1}\)) would be required to react with 5.00 g of carbon?

Carbon is the *given* substance and silicon dioxide is the *desired* substance.

\[
5.00 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} \times \frac{6 \text{ mol SiO}_2}{10 \text{ mol C}} \times \frac{60.1 \text{ g SiO}_2}{1 \text{ mol SiO}_2} = 15.0 \text{ g SiO}_2
\]

b) What minimum mass of magnesium phosphate (\(M_m = 263 \text{ g·mol}^{-1}\)) would be required to produce 12.0 g of phosphorus (\(M_m = 124 \text{ g·mol}^{-1}\))?

Phosphorus is the *given* substance and magnesium phosphate is the *desired* substance.

\[
12.0 \text{ g P}_4 \times \frac{1 \text{ mol P}_4}{124 \text{ g P}_4} \times \frac{2 \text{ mol Mg}_3(PO_4)_2}{1 \text{ mol P}_4} \times \frac{263 \text{ g Mg}_3(PO_4)_2}{1 \text{ mol Mg}_3(PO_4)_2} = 50.9 \text{ g Mg}_3(PO_4)_2
\]

c) What volume (in L) of carbon monoxide measured at 1.00 atm and 20. °C* would result if 15.0 g of carbon reacted?

Carbon is the *given* substance and carbon monoxide is the *desired* substance. In this example, we will have to use the ideal gas law (\(PV = nRT\)) to convert from moles of CO into liters of CO. We begin by calculating the number of moles of CO that should form.

\[
15.0 \text{ g C} \times \frac{1 \text{ mol C}}{12.0 \text{ g C}} \times \frac{10 \text{ mol CO}}{12.0 \text{ g C}} = 1.25 \text{ mol CO}
\]

We can now solve the ideal gas law (Eq. 1.1) for the volume. Remember that the temperature must be expressed on the Kelvin scale.

\[
T = 20. °C + 273 = 293 \text{ K}
\]

\[
V = \frac{nRT}{P} = \frac{(1.25 \text{ mol CO})(0.0821 \text{ L·atm·K}^{-1}·\text{mol}^{-1})(293 \text{ K})}{1.00 \text{ atm}} = 30.1 \text{ L}
\]

* The decimal after the 20 is used to indicate that the zero is significant. See Appendix A for more on significant figures.
LIMITING REACTANTS

In the calculations done to this point, the mass of only one substance has been given but normally several reactants are added, and it is seldom that they are all added in the exact mass ratios required for complete reaction. In this case, one reactant limits the amount of reaction because when any one reactant is completely consumed, the reaction stops. The reactant that is completely consumed is called the limiting reactant. Any reactants that are not completely consumed are excess reactants. For example, your automobile engine is powered by the combustion of gasoline, so it requires both gasoline and oxygen to run. There is an excess of oxygen in the air, so the gasoline is the limiting reactant. When the limiting reactant is consumed the reaction stops; i.e., you run out of gas even though there is plenty of oxygen still available.

All calculations pertaining to a reaction are based on its limiting reactant, so if the amount of more than one reactant is given, we must determine which is the limiting amount before we can do anything else. To determine the limiting reactant, we must determine which reactant (A) produces the least amount of product (B), which would be done as follows:

\[
\text{mol A} \times \frac{\text{coefficient of B}}{\text{coefficient of A}} = \text{mol B}
\]

The reactant that can produce the smallest number of moles of B is the limiting reactant. Thus, one way to determine the limiting reactant is to determine how much of one product each of the reactants can produce. However, we can shorten the step a little by rearranging the above to the form given in Equation 1.4

\[
\left( \frac{\text{mol A}}{\text{coefficient of A}} \right) \times \text{coefficient of B} = \text{mol B} \quad \text{Eq. 1.4}
\]

The amount of any product B that can be produced by a given number of moles of any reactant can be obtained by multiplying the (mole/coefficient) ratio of that reactant by the coefficient of the product in the balanced chemical equation. We conclude that the smallest number of moles of product is obtained from the reactant that has the smallest (mole/coefficient) ratio. Thus, the limiting reactant can be found by doing the following:

Divide the number of moles of each reactant present by its coefficient in the balanced equation. The reactant with the smallest ratio is the limiting reactant.

PRACTICE EXAMPLE 1.10

a) What mass of Ba₃(PO₄)₂ can be produced from 16.8 g of Ba(OH)₂? See Practice Example 1.9 for the chemical equation.

- Molar mass of Ba₃(PO₄)₂ = __________ g mol⁻¹
- Molar mass of Ba(OH)₂ = __________ g mol⁻¹
- Moles of Ba(OH)₂ reacting = __________ mol
- Moles of Ba₃(PO₄)₂ produced = __________ mol
- Mass of Ba₃(PO₄)₂ produced = __________ g

b) What mass of H₃PO₄ is required to react with 16.8 g of Ba(OH)₂?

- Mass = __________ g H₃PO₄
As an example, let us determine the limiting reactant when 10.0 g of C, 20.0 g of SiO₂, and 40.0 g of Mg₃(PO₄)₂ are used in the following reaction:

\[ 10C + 6SiO₂ + 2Mg₃(PO₄)₂ \rightarrow P₄ + 6MgSiO₃ + 10CO \]

First, determine the number of moles of each reactant.

- 10.0 g C \( \times \frac{1 \text{ mol C}}{12.0 \text{ g C}} \) = 0.833 mol C
- 20.0 g SiO₂ \( \times \frac{1 \text{ mol SiO₂}}{60.1 \text{ g SiO₂}} \) = 0.333 mol SiO₂
- 40.0 g Mg₃(PO₄)₂ \( \times \frac{1 \text{ mol Mg₃(PO₄)₂}}{263 \text{ g Mg₃(PO₄)₂}} \) = 0.152 mol Mg₃(PO₄)₂

Next, use the moles determined above and the coefficients in the chemical equation to determine the mole/coefficient ratios of the reactants.

\[ \frac{0.833 \text{ mol C}}{10 \text{ mol C}} = 0.0833 \quad \frac{0.333 \text{ mol SiO₂}}{6 \text{ mol SiO₂}} = 0.0555 \quad \frac{0.152 \text{ mol Mg₃(PO₄)₂}}{2 \text{ mol Mg₃(PO₄)₂}} = 0.0760 \]

SiO₂ has the smallest ratio, so it is the limiting reactant. There are two important points to make here.

1. The reactant with the smallest mass is not necessarily the limiting reactant. In this example, carbon had the smallest mass yet it has the largest mole/coefficient ratio.
2. The reactant present in the smallest number of moles is not necessarily the limiting reactant. The number of moles of magnesium phosphate present was less than half the number of moles of silicon dioxide, but SiO₂ has the smaller mole/coefficient ratio because the reaction requires three moles of SiO₂ for every one of Mg₃(PO₄)₂.

The limiting reactant in this problem is SiO₂, so all further calculations are based on the fact that all 0.333 mol SiO₂ reacts. For example, the amount of MgSiO₃ (\( M_m = 100. \text{ g-mol}^{-1} \)) that forms is determined to be

\[ 0.333 \text{ mol SiO₂} \times \frac{6 \text{ mol MgSiO₃}}{6 \text{ mol SiO₂}} \times \frac{100 \text{ g MgSiO₃}}{1 \text{ mol MgSiO₃}} = 33.3 \text{ g MgSiO₃} \]

To determine how much excess carbon there is, determine how much carbon reacts and subtract that from the amount in the original mixture.

\[ 0.333 \text{ mol SiO₂} \times \frac{10 \text{ mol C}}{6 \text{ mol SiO₂}} \times \frac{12.0 \text{ g C}}{1 \text{ mol C}} = 6.7 \text{ g C react} \]

10.0 g C initially – 6.7 g C reacts = 3.3 g C in excess

Therefore, 3.3 g C would remain unreacted in the reaction vessel because all of the SiO₂ had been consumed.
The amount of product formed depends not only on the amounts of reactants but also on the equilibrium constant for the reaction. Indeed, the problem we have just completed can be considered to be an equilibrium problem in which the equilibrium constant is very large, i.e., one in which essentially all of the limiting reactant is consumed.

We now set up the same problem in a way that leads directly to the amounts of both reactants and products present at the end of the reaction. We begin by writing the reaction and labeling three lines under it. The number of moles of each substance present before the reaction begins is placed into the first line, which is referred to as the initial line. The number of moles of each substance that reacts or forms during the reaction is placed on the following line. This line represents the changes in the amounts that result from the reaction, so it is designated as the Δ (delta) line. The Δ line is the only line to which stoichiometry is applied, and the numbers in it are based on the limiting reactant. The sum of the initial line and Δ line is placed on the final line, which represents the composition of the mixture after the reaction is complete. Together, the three lines constitute the reaction table for the reaction.

The reaction table is started by placing the initial number of moles of each substance directly under the substance in the chemical equation.

\[
\begin{align*}
10C & \quad + \quad 6SiO_2 & \quad + \quad 2Mg_3(PO_4)_2 & \rightarrow & \quad P_4 & \quad + \quad 6MgSiO_3 & \quad + \quad 10CO \\
\text{Initial} & \quad 0.833 & \quad 0.333 & \quad 0.152 & \quad 0 & \quad 0 & \quad 0 \quad \text{mol}
\end{align*}
\]

The Δ line represents the amounts that react or form during the reaction, so all values on it depend upon the limiting reactant. SiO\textsubscript{2} is the limiting reactant, so the entire amount of SiO\textsubscript{2} is entered on the Δ line. Materials that react disappear during reaction, so the amount on the Δ line under reacting substances is negative.

\[
\begin{align*}
10C & \quad + \quad 6SiO_2 & \quad + \quad 2Mg_3(PO_4)_2 & \rightarrow & \quad P_4 & \quad + \quad 6MgSiO_3 & \quad + \quad 10CO \\
\text{initial} & \quad 0.833 & \quad 0.333 & \quad 0.152 & \quad 0 & \quad 0 & \quad 0 \quad \text{mol}
\]
\]

\[
\begin{align*}
\Delta & \quad -0.333 \quad \text{mol}
\end{align*}
\]

The Δ line is completed by applying the various stoichiometric ratios to the amount of the limiting reactant that is consumed. The reactants are all disappearing and the products are all forming. Consequently, all entries on the reactant side are negative and all entries on the product side are positive. For example, consider the entry under Mg\textsubscript{3}(PO\textsubscript{4})\textsubscript{2}.

\[0.333 \text{ mol } SiO_2 \times \frac{2 \text{ mol } Mg_3(PO_4)_2}{6 \text{ mol } SiO_2} = 0.111 \text{ mol } Mg_3(PO_4)_2 \text{ reacts}\]

Mg\textsubscript{3}(PO\textsubscript{4})\textsubscript{2} reacts, so its entry on the Δ line is -0.111 mol. MgSiO\textsubscript{3} forms as the SiO\textsubscript{2}
disappears, so its entry would be determined as

\[ 0.333 \text{ mol SiO}_2 \times \frac{6 \text{ mol MgSiO}_3}{6 \text{ mol SiO}_2} = 0.333 \text{ mol MgSiO}_3 \text{ forms} \]

MgSiO₃ is produced, so its entry on the Δ line is +0.333 mol. Thus, the Δ line is

\[ 10\text{C} + 6\text{SiO}_2 + 2\text{Mg}_3(\text{PO}_4)_2 \rightarrow \text{P}_4 + 6\text{MgSiO}_3 + 10\text{CO} \]
\[ \Delta \ -0.555 \ -0.333 \ -0.111 \ +0.055 \ +0.333 \ +0.555 \text{ mol} \]

To determine the final composition, the initial and Δ lines are added as shown below.

\[
\begin{array}{cccccc}
\text{initial} & 0.833 & 0.333 & 0.152 & 0 & 0 & 0 \\
\Delta & -0.555 & -0.333 & -0.111 & +0.055 & +0.333 & +0.555 \\
\text{final} & 0.278 & 0.000 & 0.041 & 0.055 & 0.333 & 0.555 \\
\end{array}
\]

Once the reaction table is complete, we know the amounts of all of the substances present at the end of the reaction. If all of the SiO₂ reacts (zero on the final line), then there would be 0.278 mol C and 0.041 mol of Mg₃(PO₄)₂ remaining while 0.055 mol P₄, 0.333 mol MgSiO₃ and 0.555 mol CO would form.

The calculated amount of product is called the theoretical yield because it is the amount that should be produced. However, the actual yield of most reactions is less than the theoretical yield for several reasons:

- Many reactions reach an equilibrium in which a measurable amount of the limiting reactant remains.
- Purification results in some loss of product.
- Reactants are often involved in more than one type of reaction, and these side reactions can result in several different products, which reduces the quantity of the desired product.

In order to represent the efficiency of a procedure, chemists normally report the percent yield for the product. The percent yield is defined as the fraction of the theoretical yield that is actually obtained expressed as a percent.

\[
\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \quad \text{Eq. 1.5}
\]

**PRACTICE EXAMPLE 1.11**

What is the percent yield in the precipitation described in Practice Example 1.10a if 18.3 g of Ba₃(PO₄)₂ is isolated?

\[
\text{% yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \text{__________ %}
\]
Example 1.14

a) Aspirin (acetylsalicylic acid, C₉H₈O₄) is prepared from the following reaction of salicylic acid (C₇H₆O₃) and acetic anhydride (C₄H₆O₃). The other product is acetic acid (C₂H₄O₂).

\[
\text{C}_7\text{H}_6\text{O}_3 + \text{C}_4\text{H}_6\text{O}_3 \rightarrow \text{C}_9\text{H}_8\text{O}_4 + \text{C}_2\text{H}_4\text{O}_2
\]

If 20.00 g of salicylic acid and 17.00 g of acetic anhydride react, what would be the composition of the reaction mixture at completion if the reaction occurs with a 100% yield?

First, convert the initial masses into moles by dividing by the molar masses.

\[
\begin{align*}
20.00 \text{ g} \text{C}_7\text{H}_6\text{O}_3 &= 0.1448 \text{ mol C}_7\text{H}_6\text{O}_3 & 138.12 \text{ g mol}^{-1} \\
17.00 \text{ g} \text{C}_4\text{H}_6\text{O}_3 &= 0.1665 \text{ mol C}_4\text{H}_6\text{O}_3 & 102.09 \text{ g mol}^{-1}
\end{align*}
\]

There are no products initially, so the initial line of the reaction table is

<table>
<thead>
<tr>
<th></th>
<th>C₇H₆O₃</th>
<th>C₄H₆O₃</th>
<th>C₉H₈O₄</th>
<th>C₂H₄O₂</th>
<th>Initial</th>
<th>0.1448</th>
<th>0.1665</th>
<th>0</th>
<th>0</th>
<th>mol</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Δ</td>
<td>-0.1448</td>
<td>-0.1448</td>
<td>+0.1448</td>
<td>+0.1448</td>
<td>mol</td>
</tr>
<tr>
<td>Final</td>
<td>0</td>
<td>0.0217</td>
<td>0.1448</td>
<td>0.1448</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Multiplying each of the moles by the molar mass, we obtain the following masses,

\[
\begin{align*}
(0.0217 \text{ mol C}_7\text{H}_6\text{O}_3)(102.09 \text{ g mol}^{-1}) &= 2.22 \text{ g C}_7\text{H}_6\text{O}_3 \\
(0.1448 \text{ mol C}_9\text{H}_8\text{O}_4)(180.16 \text{ g mol}^{-1}) &= 26.09 \text{ g C}_9\text{H}_8\text{O}_4 \\
(0.1448 \text{ mol C}_2\text{H}_4\text{O}_2)(60.05 \text{ g mol}^{-1}) &= 8.70 \text{ g C}_2\text{H}_4\text{O}_2
\end{align*}
\]

As a check, we note that the total mass at the completion of the reaction is the same as the total starting mass.

b) If the reaction produces an 85.3% yield, what mass of aspirin would be isolated?

The theoretical yield of aspirin is 26.09 g, but only 85.3% of this mass was isolated. Therefore, the actual yield is (0.853)(26.09) = 22.3 g.

---

PRACTICE EXAMPLE 1.12

The following is the first step of Ostwald process, a three-step reaction sequence that converts NH₃ into HNO₃:

\[
4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)
\]

10.00 g each of NH₃ and O₂ are allowed to react. Assume complete reaction and determine the contents of the flask after the reaction is complete.

<table>
<thead>
<tr>
<th>moles of NH₃</th>
<th>= ________ mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>moles of O₂</td>
<td>= ________ mol</td>
</tr>
</tbody>
</table>

Determine how much NO could be produced by each

<table>
<thead>
<tr>
<th>from NH₃</th>
<th>= ________ mol NO</th>
</tr>
</thead>
<tbody>
<tr>
<td>from O₂</td>
<td>= ________ mol NO</td>
</tr>
</tbody>
</table>

limiting reactant is ________

| = ________ mol H₂O form |
| = ________ mol O₂ react |
| = ________ mol NH₃ react |

Reaction table

\[
\begin{align*}
4\text{NH}_3(g) + 5\text{O}_2(g) & \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g) \\
\text{initial} & \Delta \\
\text{final} & \\
\end{align*}
\]

Final contents of flask

| _____ g NH₃ |
| ______ g O₂ |
| ______ g NO |
| ______ g H₂O |
1.6 CHAPTER SUMMARY AND OBJECTIVES

The elemental composition of a compound is usually defined in terms of the mass ratios of its elements, while the chemical formula expresses the mole ratios of its elements. It is important to be able to convert between these two methods of describing composition. The elemental composition of a substance can be determined from its chemical formula by dividing the mass contribution of each atom in the formula by the molar mass of the substance. The resulting fraction is usually expressed as a percent. The formula of a compound can be determined by performing the reverse procedure - the mass or percent of each element present is converted to moles. The moles of the elements are then expressed as simple whole number ratios to obtain the formula. The resulting formula is the simplest or empirical formula because it expresses only the simplest ratios of the atoms present. The molecular formula can be determined from the empirical formula and the molar mass.

The number of moles of a given species can be converted to the chemically equivalent amount of a sought species with the stoichiometric ratio. Stoichiometric ratios are derived from the subscripts in a formula or the coefficients in a balanced equation. Many chemical equations can be balanced by inspection by fixing the coefficient of the substance with the greatest number of atoms and then balancing each atom. The stoichiometric amount of any desired substance can then be determined from an amount of any given substance as follows:

\[
\frac{\text{moles given}}{\text{moles given}} \times \frac{\text{moles sought}}{\text{moles given}} = \text{moles sought}
\]

The amount of the limiting reactant in a chemical reaction dictates the amount of product that can form. Other reactants in the reaction are said to be in excess. The limiting reactant is that reactant with the smallest mole/coefficient ratio. All further calculations are based on the limiting reactant. The amounts of all products and excess reactants that remain after the complete reaction of the limiting reactant are most easily calculated by using a reaction table. The initial line consists of the number of moles of each substance present at the start of the reaction, the \( \Delta \) line is composed of the number of moles of each substance that is produced or consumed, and the final line is the sum of the initial and \( \Delta \) lines. The percent yield of a reaction is defined as the ratio of the actual yield to the theoretical yield expressed as a percent.

ANSWERS TO PRACTICE EXAMPLES

1.1 a) 88.0 g·mol\(^{-1}\); 253 g
   b) 60.0 g·mol\(^{-1}\); 2.69 g
   c) 174 g·mol\(^{-1}\); 0.107 mol
   d) 98.0 g·mol\(^{-1}\); 2.72 mmol

1.2 a) 1.60 \times 10^{21} \text{ molecules of CO}_2
   b) 6.69 \times 10^{16} \text{ molecules of N}_2\text{O}_5

1.3 1.433 atm; 348 K; 0.0251 mol; 1.10 g

1.4 69.07% Ba, 14.8% Co, 16.1% O

1.5 206.98 g/mol, 19.36% Ca, 34.26% Cl, 46.38% O

1.6 41.1% H\(_2\)O, 2.24 mmol CoSO\(_4\), 13.4 mmol H\(_2\)O, 
\[ x = 6, \text{CoSO}_4\cdot6\text{H}_2\text{O} \]

1.7 0.0178 mol C; 0.0356 mol H; 2 mol H/1 mol C; 
   empirical: CH\(_2\); molecular: C\(_6\)H\(_{12}\)

1.8 26.6 g Ba

1.9 3\text{Ba(OH)}_2 + 2\text{H}_3\text{PO}_4 \rightarrow \text{Ba}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}

1.10 \text{Ba(OH)}_2: 171.31 g/mol; 0.0981 mol react
   \text{Ba}_3(\text{PO}_4)_2: 601.93 g/mol; 19.7 g produced
   \text{H}_3\text{PO}_4: 97.99 g/mol; 6.4 g required

1.11 92.9%

1.12 limiting reactant: O\(_2\)
   Final contents: 5.74 g NH\(_3\); no O\(_2\); 7.50 g NO; 6.76 g H\(_2\)O
After studying the material presented in this chapter and the relevant appendices, you should be able to:

1. convert between mass and moles (Section 1.1);
2. convert between mass or moles and numbers of atoms or molecules (Section 1.1);
3. determine moles of gas from P, V, and T (Section 1.1);
4. determine percent composition of a compound (Section 1.2);
5. determine the simplest formula of a compound from the relative amounts of each of the elements present in a sample (Section 1.2);
6. determine a molecular formula from the simplest formula and molar mass (Section 1.2);
7. use the chemical formula of a compound to determine the mass of one element that is combined with a given mass of another element in the compound (Section 1.3);
8. balance a chemical equation by inspection (Section 1.4);
9. write the stoichiometric ratio relating two substances involved in a chemical reaction (Section 1.5);
10. convert the mass of any substance in a reaction into the stoichiometrically equivalent mass of any other substance involved in the reaction (Section 1.5);
11. determine the limiting reactant of a reaction (Section 1.5);
12. combine the ideal gas law and stoichiometry to determine the amount of gas formed or consumed in a reaction (Section 1.5);
13. determine the complete composition of a reaction mixture after the reaction is complete (Section 1.5); and
14. determine the percent yield given the actual yield of a reaction or use the percent yield to determine the actual yield (Section 1.5).
1.7 EXERCISES

MASS AND MOLES

1. Determine molar masses for the following:
   a) C_{22}H_{10}O_{2}    b) Ca(NO_{3})_{2}    c) P_{2}O_{5}    d) Al_{2}(CO_{3})_{3}

2. Determine molar masses for the following:
   a) Mg(C_{2}H_{3}O_{2})_{2}    b) PtCl_{2}(NH_{3})_{2}    c) (NH_{4})_{2}SO_{4}    d) C_{12}H_{27}NO_{3}

3. Determine mass of each of the following:
   a) 0.694 mol C_{22}H_{10}O_{2}    b) 2.84 mol Ca(NO_{3})_{2}    c) 0.00652 mol P_{2}O_{5}    d) 8.44 mol Al_{2}(CO_{3})_{3}

4. Determine mass of each of the following:
   a) 1.86 mol Mg(C_{2}H_{3}O_{2})_{2}    b) 0.0356 mol PtCl_{2}(NH_{3})_{2}    c) 18.4 mol (NH_{4})_{2}SO_{4}    d) 4.88 mol C_{12}H_{27}NO_{3}

5. Determine mass of each of the following:
   a) 2.24x10^{20} molecules of CO_{2}    b) 2.24x10^{24} molecules of H_{2}    c) 12 C atoms    d) 8.66x10^{18} Pt atoms

6. Determine mass of each of the following:
   a) 2.00 million ammonia molecules    b) 1 water molecule    c) 6.02x10^{23} PF_{3} molecules    d) 4.02x10^{28} C atoms

7. How many moles of people were on the earth when the population was 6.3 billion (6.3x10^{9}) people?

8. It is estimated that there are over 400 billion (4x10^{11}) stars in the Milky Way galaxy. How many moles of stars is that?

9. A bottle contains 12.6 g of (NH_{4})_{3}PO_{4}.
   a) How many moles of (NH_{4})_{3}PO_{4} does it contain?
   b) How many oxygen atoms does it contain?
   c) What mass of nitrogen atoms does it contain?
   d) How many moles of H does it contain?

10. A heaping teaspoon of sugar (C_{12}H_{22}O_{11}) has a mass of 8.0 g.
    a) How many moles of sugar does it contain?
    b) How many oxygen atoms does it contain?
    c) What mass of carbon atoms does it contain?
    d) How many moles of H does it contain?

SUBSTANCE STOICHIOMETRY

11. What is the simplest formula of each of the following compounds?
    a) C_{22}H_{10}O_{2}    b) C_{3}H_{6}O    c) C_{6}H_{6}    d) C_{3}H_{6}O_{3}

12. What is the simplest formula of each of the following compounds?
    a) Na_{2}S_{2}O_{8}    b) B_{2}H_{6}    c) N_{3}S_{3}Cl_{3}    d) Na_{2}Re_{2}Cl_{8}

13. What is the elemental composition of each of the molecules in Exercise 11? Express your answer as mass percents?

14. What is the elemental composition of each of the compounds in Exercise 12, Express your answer as mass percents?

15. How many moles of magnesium are present in a sample of each of the following that contains 3.0 moles of oxygen atoms?
    a) MgSO_{4}    b) MgSO_{3}    c) Mg_{3}(PO_{4})_{2}    d) Mg(ClO_{3})_{2}

16. How many grams of nitrogen atoms are present in a sample of each of the following that contains 1.25 moles of oxygen atoms?
    a) NO    b) MgSO_{3}    c) N_{2}O_{5}    d) NH_{4}NO_{3}

17. What mass of Al is in a sample of Al_{2}(SO_{4})_{3} that contains 3.2 grams of S?

18. What mass of potassium is in a sample of potassium carbonate (K_{2}CO_{3}) that contains 12.0 g of oxygen?

19. Caffeine has the molecular formula C_{8}H_{10}N_{4}O_{2}. What mass of caffeine contains 5.0 mg of nitrogen?

20. Nicotine is C_{10}H_{14}N_{2}. What mass of nicotine contains 1.5 moles of nitrogen?

21. What mass of Na_{2}CO_{3} contains 2.1x10^{22} oxygen atoms?

22. What mass of oxygen is in a sample of Na_{3}PO_{4} that contains 3.5x10^{21} sodium atoms?

23. What mass of KCl was in a solution if all of the chloride in the solution was precipitated as 1.68 g of PbCl_{2}?

24. What mass of Ag_{2}SO_{4} was in a solution if all of the silver was precipitated as 375 mg of Ag_{3}PO_{4}?

DETERMINING CHEMICAL FORMULAS

25. What is the simplest formula of a compound in which 0.362 mol X is combined with 1.267 mol Y? How many moles of X are present in 6.336 mol of the compound?
26. What is the simplest formula of a compound if a sample of the compound contains 0.236 mol X, 0.354 mol Y, and 0.590 mol Z? How many moles of Z would be in a sample that contained 0.668 mol X?

27. What is the simplest formula of a hydrocarbon that is 81.71% C?

28. What is the empirical formula of a rhenium oxide that is 76.88% Re?

29. Ibuprofen (Advil® or Motrin®) is an anti-inflammatory agent that is 75.69% C, 6.80% H and 15.51% O. What is the simplest formula of ibuprofen?

30. Acetaminophen (Tylenol®) is an analgesic (pain killer) and an antipyretic (fever reducer) that is 63.56% C, 6.00% H, 9.27% N and 21.17% O. What is the empirical formula of acetaminophen?

31. The sugar arabinose, found in ripe fruits, is 40.00% C, 6.71% H and 53.29% O and has a molar mass of 150. g·mol⁻¹. What is the molecular formula for this compound?

32. What is the simplest formula of a compound that is 39.81% Cu, 20.09% S, and 40.10% O?

33. Burning 1.346 g of chromium in air results in 1.967 g of an oxide. What is the simplest formula of the oxide of chromium?

34. A 3.228-g sample of platinum oxide is found to contain 2.773 g of platinum. What is the empirical formula of the oxide?

35. A 2.500-g sample of an oxide of lead produces 0.376 g of water when reduced with hydrogen. What is the simplest formula of this lead oxide? Assume all of the oxygen in the oxide is converted to water.

36. What is the empirical formula of a hydrocarbon if combustion of 1.00 mg produces 3.14 mg of CO₂ and 1.29 mg of H₂O? If its molar mass is around 40 g·mol⁻¹, what is its molecular formula?

37. A 0.540-g sample of Anavenol, a compound containing C, H, and O that is used as an anesthetic in veterinary surgeries is analyzed by combustion. What is the empirical formula if the combustion produces 0.310 g of H₂O and 1.515 g of CO₂? If its molar mass is 188.22 g·mol⁻¹, what is the molecular formula for Anavenol?

38. Antifreeze (ethylene glycol) contains carbon, hydrogen and oxygen. Combustion of 50.00 mg of ethylene glycol yields 43.55 mg of H₂O and 70.97 mg of CO₂. What is the empirical formula of ethylene glycol? The molar mass of ethylene glycol is 62.0 g·mol⁻¹, what is its molecular formula?

39. KClO₃ produces KCl and O₂ upon heating. What is the value of x if a 22.6-g sample produces 7.07 L of O₂ at 0.956 atm and 25 °C?

40. A 0.525-g sample of an iron carbonyl, Fe(CO)₅, is heated to remove all of the CO. The CO gas is collected in a 0.500-L flask at 26 °C. What is the empirical formula of the carbonyl if the pressure of the CO is 499 torr?

41. Heating a 27.7-mg sample of MnSO₄·xH₂O results in 15.1 mg of anhydrous MnSO₄. What is value of x?

42. What mass of MgCO₃ contains the same mass of oxygen as does 376 mg of MgCr₂O₇?

BALANCING EQUATIONS

43. Balance the equations by inspection:
   a) __Al₂S₃ + ___H₂O → __Al(OH)₃ + ___H₂S
   b) __Fe₂O₃ + ___H₂ → __Fe + ___H₂O
   c) __Al + ___H₂SO₄ → __Al₂(SO₄)₃ + ___H₂
   d) __CH₃OH + ___O₂ → __CO₂ + ___H₂O
   e) __KOH + ___H₃PO₄ → __K₂HPO₄ + ___H₂O
   f) __Ag + ___H₂S + ___O₂ → __Ag₂S + ___H₂O

44. Balance the equations by inspection:
   a) __P₂O₅ + ___H₂O → __H₃PO₃
   b) __NaOH + ___HCl → __HOCl + ___N₂ + ___NaCl
   c) __H₃PO₄ + ___NH₃ → __(NH₄)₂HPO₄
   d) __Bi₂O₃ + ___C → __Bi + ___CO
   e) __HCl + ___MnO₂ → __MnCl₂ + ___H₂O + ___Cl₂
   f) __Ca₃N₂ + ___H₂O → __Ca(OH)₂ + ___NH₃

45. Balance the equations by inspection:
   a) __FeS₂ + ___O₂ → __FeSO₄ + ___SO₃
   b) __S₂Cl₂ + ___H₂O → __SO₂ + __HCl + ___S
   c) __V₂O₅ + ___C + ___Cl₂ → __VOCl₃ + ___COCl₂
   d) __NH₃ + ___O₂ → ___N₂O + ___H₂O
   e) __BF₃ + ___NaBH₄ → __NaBF₄ + ___B₂H₆
   f) __Al₄C₃ + ___HCl → __CH₄ + ___AlCl₃

46. Balance the equations by inspection:
   a) __Cr₂O₃ + ___H₂O → __Cr(OH)₃
   b) __PCl₅ + ___H₂O → __H₃PO₄ + ___HCl
   c) __Mg₂C + ___H₂O → __Mg(OH)₂ + ___CH₄
d) $\_\_\text{BF}_3 + \_\_\text{NaH} \rightarrow \_\_\text{NaBF}_4 + \_\_\text{B}_2\text{H}_6$

e) $\_\_\text{SiO}_2 + \_\_\text{Ca}_3(\text{PO}_4)_2 \rightarrow \_\_\text{P}_2\text{O}_5 + \_\_\text{CaSiO}_3$

f) $\_\_\text{Ba}_3\text{P}_2 + \_\_\text{H}_2\text{O} \rightarrow \_\_\text{Ba(OH)}_2 + \_\_\text{PH}_3$

**REACTION STOICHIOMETRY**

47. A mixture of 3.0 mol of CS$_2$ and 2.0 mol of O$_2$ reacts according to the equation: $\text{CS}_2 + 3\text{O}_2 \rightarrow \text{CO}_2 + 2\text{SO}_2$

a) What is the limiting reactant?

b) How many moles of SO$_2$ are produced?

c) How many moles of which reactant are unreacted?

d) If 72 g of SO$_2$ are actually isolated, what is the percent yield?

48. An excess of O$_2$ is added to 4.86 g of Fe and allowed to react. What is the percent yield if 6.76 g of Fe$_2$O$_3$ are isolated?

49. Consider the reaction $\text{N}_2\text{O}_4 + 2\text{N}_2\text{H}_4 \rightarrow 3\text{N}_2 + 4\text{H}_2\text{O}$

a) How many moles of N$_2$ are formed by reaction of 5.0 g of N$_2$H$_4$?

b) What mass of N$_2$O$_4$ would be required for Part a?

c) What is the percent yield if 4.8 g of water is produced?

50. What mass of oxygen is required for the complete combustion of 7.65 g of propane (C$_3$H$_8$) to produce CO$_2$ and H$_2$O?

51. What mass of HCl is produced by the reaction of 23.6 g of PCl$_3$ and water? The other product is H$_3$PO$_3$.

52. How many liters of O$_2$ gas measured at 835 torr and 250 °C are formed by the decomposition of 236 g of KClO$_3$? The other product is KCl.

53. Consider the following reaction that occurs at 1000 °C: $4\text{NH}_3(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{N}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$.

A mixture of 2.65 atm of NH$_3$ and 3.80 atm of O$_2$ reacts to completion (no limiting reactant remains). Determine the pressures of all gases remaining when the reaction is complete. (Hint: because the reaction is carried out at constant temperature and volume, the pressures are proportional to the number of moles.) What is the total pressure inside the vessel at the beginning and end of the reaction? Why are the total pressures different?

54. Construct the reaction table for the reaction of 2.0 mol Fe$_3$O$_4$ and 6.0 mol H$_2$ to produce elemental iron and water. How many moles of iron form and how many moles of the excess reactant are unused?

55. Construct the reaction table for the reaction of 7.0 g of N$_2$ and 6.0 g of H$_2$ to form ammonia. What mass of ammonia forms and what mass of the excess reactant remains after reaction?

56. Consider the reaction between H$_2$PO$_4$ and NH$_3$ to produce (NH$_4$)$_2$HPO$_4$.

a) What mass of ammonia would have to be added to 20.0 g of phosphoric acid if a 10% excess of ammonia is required?

b) What is the theoretical yield of ammonium hydrogenphosphate under the conditions given in Part a?

57. Construct a reaction table for the reaction of 0.200 mol of iron(III) oxide and 0.270 mol of carbon to produce elemental iron and carbon monoxide. What is the percent yield if 19.4 g of iron are produced?

58. Construct the reaction table for the reaction of 1.46 mol Al and 3.61 mol HCl to produce AlCl$_3$ and H$_2$.

59. The most common acid in acid rain is sulfuric acid (H$_2$SO$_4$). When sulfuric acid reacts with sodium hydroxide (NaOH), sodium sulfate is formed along with water. The reaction is $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$

A 10.0-L sample of rain water was treated with a 0.200-g tablet of NaOH. When the reaction was complete, 0.0018 moles of NaOH remained unreacted.

a) What was the limiting reagent in this reaction?

b) How many grams of H$_2$SO$_4$ were in the 10.0-L sample of rain water?

c) How many moles of H$_2$SO$_4$ were present in each liter of rain water?

60. One step in the production of margarine from vegetable oils is the hydrogenation of the double bonds. In an experiment to test a new hydrogenation catalyst, a 2.36-g sample of linolenic acid, C$_{18}$H$_{30}$O$_2$, was placed in a sealed flask along with a catalyst. Then 1.50 L of H$_2$ gas (measured at 1.0 atm pressure and 25 °C) was bubbled into the mixture. The completely hydrogenated product of the reaction is stearic acid, C$_{18}$H$_{36}$O$_2$.

a) Write the balanced reaction for the complete hydrogenation of linolenic acid to stearic acid.

b) What was the limiting reagent in this reaction mixture?

c) After the reaction, 2.06 g of stearic acid was recovered. What was the percent yield of the reaction?
61. It is desired to remove the lead from a solution containing 6.41 g of Pb(NO_3)_2 by adding KCl and precipitating PbCl_2. What mass of KCl should be added if a 15.0% excess is required? What mass of PbCl_2 would form? The other product is KNO_3.

62. Consider the reaction 5P_4O_6 + 8I_2 → 4P_2I_4 + 3P_4O_{10}.
   a) How many grams of I_2 should be added to 4.50 g of P_4O_6 in order to have a 10.0% excess?
   b) What is the theoretical yield of P_2I_4?
   c) How many grams of P_2I_4 would be isolated if actual yield is 83.7%?

63. Consider the reaction of 27.8 g of FeS_2 with O_2 to produce Fe_2O_3 and SO_2.
   a) What mass of oxygen would be required for a 20% excess?
   b) What is the theoretical yield of Fe_2O_3?
   c) What mass of SO_2 would form if actual yield is 94.2%?

64. Construct the reaction table for the reaction of 19.25 g V_2O_5, 12.80 g C, and 30.66 g of Cl_2 to produce COCl_2 and VOCl_3.

MISCELLANEOUS PROBLEMS

65. What element forms an oxide X_2O_3 that is 88.39% X by mass?

66. The compound X_2Y_3 is found to be 75.0% X. What is the ratio of the molar masses?

67. Aspartame, C_14H_18N_2O_5, is the active ingredient in Nutrasweet®.
   a) What is the elemental composition of aspartame expressed as percents?
   b) What is the mass of a sample of aspartame that contains 2.6 mg of carbon?
   c) A tablet of Equal® has a mass of 0.088 g and the “sweetness of one teaspoon of sugar.” A teaspoon of sugar (C_12H_22O_11) has a mass of 4.8 g. Assume the Equal® tablet is 30% aspartame and estimate the relative “sweetness” of a molecule of aspartame and a molecule of sugar.

68. Methyl alcohol, CH_3OH, is a clean-burning fuel. It can be synthesized from CO(g) and H_2(g), obtained from coal and water, respectively. If you start with 12.0 g of H_2 and 74.5 g of CO, what mass of methyl alcohol can be obtained theoretically?

69. Cisplatin, Pt(NH_3)_2Cl_2, a compound used in chemotherapy for cancer patients, is synthesized by reacting ammonia with tetrachloroplatinate, K_2PtCl_4, to form the product and potassium chloride.
   a) What is the maximum mass of cisplatin that can be formed by the reaction of 60.0 g of K_2PtCl_4 and 40.0 g of ammonia?
   b) What is the percent yield if 35.0 g are obtained experimentally?

70. Excess hydrochloric acid reacts with 0.750 g of aluminum to form aluminum chloride and hydrogen gas.
   a) How many liters of gas would be collected at 0 °C and 1.00 atm?
   b) How many grams of aluminum chloride would be formed?

71. A 5.00-g mixture of NaCl and BaCl_2 is dissolved in water, then a solution of Na_2SO_4 is added to precipitate BaSO_4. What percent of the mass of the original mixture is due to BaCl_2 if the mass of BaSO_4 is 2.78 g?

72. What is the molar mass of hemoglobin if its four iron atoms are 0.33% of its mass?

73. Chlorophyll contains 2.72% magnesium. If there is one magnesium per chlorophyll molecule, what is the molar mass of chlorophyll?

74. Vitamin B1 is 16.6% N by mass and contains 4 nitrogen atoms. What is its molar mass?

75. A mixture of NH_4Cl and NH_4Br is 27.4% NH_4Cl by mass. What mass of the mixture contains 0.200 mol NH_4^+ ions?

76. A mixture is 18.6% NaCl, 22.1% CaCl_2, and 59.3% NaBr.
   a) What mass of the mixture contains 0.500 mole of chloride ions?
   b) How many moles of sodium ions are present in 23.8 g of the mixture?

77. A metal (M) reacts with acid according to the following equation:
   2M + 6HCl → 2MCl_3 + 3H_2.
   What is the metal if reaction of 0.305 g of M produces 161 mL of H_2 measured at 23 °C and 753 torr?

78. What is the identity of a metal (M) if 4.26 g of MCl_2 produces 11.00 g of AgCl upon reaction with excess AgNO_3? The balanced equation is
   MCl_2 + 2AgNO_3 → 2AgCl + M(NO_3)_2.

79. Epsom salts have the formula MgSO_4•xH_2O. What is the value of x if drying a 3.268-g sample results in 1.596 g of anhydrous MgSO_4?
80. The inflation in automotive air bags is the result of the rapid decomposition of sodium azide (a compound that contains only Na and N) to metallic sodium and nitrogen gas. What is the simplest formula of sodium azide if the decomposition 8.462 g of sodium azide produces 4.8052 L of N₂ measured at 23.6 °C and 752 torr?

81. Sodium nitride is prepared by reacting nitrogen gas with sodium.

\[ 6\text{Na} + \text{N}_2 \rightarrow 2\text{Na}_3\text{N} \]

How many liters of nitrogen measured at 765 torr and 27.5 °C are required for the complete reaction of 7.22 g of Na? How many grams of sodium nitride would be produced?

82. Ethyl acetate, the active ingredient in nail polish remover, is an ester prepared by the reaction of acetic acid (vinegar) and ethanol (grain alcohol):

\[ \text{CH}_3\text{COOH} + \text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O} \]

The amount of ester is increases by removing water (LeChatelier’s principle). In a given reaction, 7.65 g CH₃COOH and 9.88 g CH₃OH are mixed and allowed to react. What is the percent yield if 8.96 g of the ester is isolated?

83. How many carbon atoms are present in a 2.0 carat diamond? 1 carat = 0.200 g.

84. The explosion of nitroglycerin is due to the following exothermic reaction:

\[ 4\text{C}_3\text{H}_5\text{N}_3\text{O}_9(l) \rightarrow 12\text{CO}_2(g) + 6\text{N}_2(g) + \text{O}_2(g) + 10\text{H}_2\text{O}(g) \]

What volume of gas is produced at 1.00 atm and 200 °C by the reaction of 2.35 g of nitroglycerin?

85. A mixture of KBr and MgBr₂, which has a mass of 6.81 g, is dissolved in water. An excess of AgNO₃ is then added to the solution to precipitate all of the bromide as AgBr. What are the mass percents of K and Mg in the mixture if 13.24 g of AgBr precipitate?

86. 3.62 g of a Group 1A metal reacts with an excess of oxygen to produce 4.36 g of its oxide. What is the metal?

87. Vanillin, which is the primary ingredient in vanilla flavoring, contains C, H, and O. What is its empirical formula if the combustion of 0.6427 g of vanillin produces 0.3043 g of H₂O and 1.487 g of CO₂? If the molar mass of vanillin is found to be near 150 g·mol⁻¹, what is its molecular formula?

88. 6.824 g of an iron chloride is dissolved in acid. Lead nitrate is then added to the solution to precipitate all of the chloride as PbCl₂. What is the empirical formula of an iron chloride if 17.568 g of PbCl₂ is produced?

89. Analysis of a compound shows that it is 17.71% N, 40.55% S, and 40.46% O by mass. It is also known to contain H. What is its empirical formula? If its molar mass is close to 240 g·mol⁻¹, what is it molecular formula?

90. It is desired to prepare exactly 5.0 g of PbCl₂ by the reaction of KCl and Pb(NO₃)₂. How many grams of each starting material should be used if a 10% excess of KCl is recommended and a 78% yield can be expected?

91. Construct a reaction table for the reaction of 12.0 g N₂ with 21.0 g O₂ to produce N₂O₅. How many g of N₂O₅ are produced, and what is the mass of the excess reactant if the reaction goes 100% to completion?