Chapter 3

Stoichiometry

- Chapter Introduction
- 3.1 Counting by Weighing
- 3.2 Atomic Masses
- 3.3 The Mole
- 3.4 Molar Mass
- 3.5 Learning to Solve Problems
- 3.6 Percent Composition of Compounds
- 3.7 Determining the Formula of a Compound
- 3.8 Chemical Equations
  - Chemical Reactions
    - The Meaning of a Chemical Equation
- 3.9 Balancing Chemical Equations
- 3.10 Stoichiometric Calculations: Amounts of Reactants and Products
- 3.11 The Concept of Limiting Reactant
  - A. Determination of Limiting Reactant Using Reactant Quantities
  - B. Determination of Limiting Reactant Using Quantities of Products Formed
- For Review
  - Key Terms
  - Summary
  - Review Questions
  - Active Learning Questions
  - Questions
  - Exercises: Atomic Masses and the Mass Spectrometer
  - Exercises: Moles and Molar Masses
  - Exercises: Percent Composition
Chapter Introduction

Fireworks provide a spectacular example of chemical reactions.
Chemical reactions have a profound effect on our lives. There are many examples: Food is converted to energy in the human body; nitrogen and hydrogen are combined to form ammonia, which is used as a fertilizer; fuels and plastics are produced from petroleum; the starch in plants is synthesized from carbon dioxide and water using energy from sunlight; human insulin is produced in laboratories by bacteria; cancer is induced in humans by substances from our environment; and so on, in a seemingly endless list. The central activity of chemistry is to understand chemical changes such as these, and the study of reactions occupies a central place in this book. We will examine why reactions occur, how fast they occur, and the specific pathways they follow.

In this chapter we will consider the quantities of materials consumed and produced in chemical reactions. This area of study is called chemical stoichiometry (the calculation of the quantities of material consumed and produced in chemical reactions.) (pronounced stoy·ki·om·′eh·tre). To understand chemical stoichiometry, you must first understand the concept of relative atomic masses.

3.1 Counting by Weighing

Suppose you work in a candy store that sells gourmet jelly beans by the bean. People come
in and ask for 50\textit{beans}, 100\textit{beans}, 1000\textit{beans}, and so on, and you have to count them out—a tedious process at best. As a good problem solver, you try to come up with a better system. It occurs to you that it might be far more efficient to buy a scale and count the jelly beans by weighing them. How can you count jelly beans by weighing them? What information about the individual beans do you need to know?

Assume that all of the jelly beans are identical and that each has a mass of 5g. If a customer asks for 1000\textit{jelly beans}, what mass of jelly beans would be required? Each bean has a mass of 5g, so you would need $1000\textit{beans} \times 5\text{g/bean}$, or 5000g (5kg). It takes just a few seconds to weigh out 5kg of jelly beans. It would take much longer to count out 1000 of them.

Jelly beans can be counted by weighing.

In reality, jelly beans are not identical. For example, let’s assume that you weigh 10\textit{beans} individually and get the following results:

<table>
<thead>
<tr>
<th>Bean</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>5.1g</td>
</tr>
<tr>
<td>2</td>
<td>5.2g</td>
</tr>
<tr>
<td>3</td>
<td>5.0g</td>
</tr>
<tr>
<td>4</td>
<td>4.8g</td>
</tr>
<tr>
<td>5</td>
<td>4.9g</td>
</tr>
<tr>
<td>6</td>
<td>5.0g</td>
</tr>
</tbody>
</table>
Can we count these nonidentical beans by weighing? Yes. The key piece of information we need is the average mass of the jelly beans. Let’s compute the average mass for our 10-bean sample.

\[
\text{Average mass} = \frac{\text{total mass of beans}}{\text{number of beans}}
\]

\[
= \frac{5.1\text{g} + 5.2\text{g} + 5.0\text{g} + 4.8\text{g} + 4.9\text{g} + 5.0\text{g} + 5.1\text{g} + 4.9\text{g} + 5.0\text{g}}{10}
\]

\[
= \frac{50.0}{10} = 5.0\text{g}
\]

The average mass of a jelly bean is 5.0g. Thus, to count out 1000 beans, we need to weigh out 5000g of beans. This sample of beans, in which the beans have an average mass of 5.0g, can be treated exactly like a sample where all of the beans are identical. Objects do not need to have identical masses to be counted by weighing. We simply need to know the average mass of the objects. For purposes of counting, the objects behave as though they were all identical, as though they each actually had the average mass.

We count atoms in exactly the same way. Because atoms are so small, we deal with samples of matter that contain huge numbers of atoms. Even if we could see the atoms, it would not be possible to count them directly. Thus we determine the number of atoms in a given sample by finding its mass. However, just as with jelly beans, to relate the mass to a number of atoms, we must know the average mass of the atoms.

Chapter 3: Stoichiometry: 3.2 Atomic Masses
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3.2 Atomic Masses

AP Learning Objectives

LO 1.14 The student is able to use data from mass spectrometry to identify the elements and the masses of individual atoms of a
As we saw in Chapter 2, the first quantitative information about atomic masses came from the work of Dalton, Gay-Lussac, Lavoisier, Avogadro, and Berzelius. By observing the proportions in which elements combine to form various compounds, nineteenth-century chemists calculated relative atomic masses. The modern system of atomic masses, instituted in 1961, is based on $^{12}\text{C}$ ("carbon twelve") as the standard. In this system, $^{12}\text{C}$ is assigned a mass of exactly 12 atomic mass units ($\text{u}$), and the masses of all other atoms are given relative to this standard.

The most accurate method currently available for comparing the masses of atoms involves the use of the **mass spectrometer** (an instrument used to determine the relative masses of atoms by the deflection of their ions on a magnetic field). In this instrument, diagramed in Fig. 3.1, atoms or molecules are passed into a beam of high-speed electrons, which knock electrons off the atoms or molecules being analyzed and change them into positive ions. An applied electric field then accelerates these ions into a magnetic field. Because an accelerating ion produces its own magnetic field, an interaction with the applied magnetic field occurs, which tends to change the path of the ion. The amount of path deflection for each ion depends on its mass—the most massive ions are deflected the smallest amount—which causes the ions to separate, as shown in Fig. 3.1. A comparison of the positions where the ions hit the detector plate gives very accurate values of their relative masses. For example, when $^{12}\text{C}$ and $^{13}\text{C}$ are analyzed in a mass spectrometer, the ratio of their masses is found to be

\[
\frac{\text{Mass } ^{13}\text{C}}{\text{Mass } ^{12}\text{C}} = 1.0836129
\]

**Figure 3.1**

(left) A scientist injecting a sample into a mass spectrometer. (right) Schematic diagram of a mass spectrometer.

Since the atomic mass unit is defined such that the mass of $^{12}\text{C}$ is exactly 12 atomic mass units, then on this same scale,
The masses of other atoms can be determined in a similar fashion.

The mass for each element is given in the table inside the front cover of this text. This value, even though it is actually a mass, is sometimes called the \textit{atomic weight} for each element.

Look at the value of the atomic mass of carbon given in this table. You might expect to see 12, since we said the system of atomic masses is based on $^{12}\text{C}$. However, the number given for carbon is not 12 but 12.01. Why? The reason for this apparent discrepancy is that the carbon found on earth (natural carbon) is a mixture of the isotopes $^{12}\text{C}$, $^{13}\text{C}$, and $^{14}\text{C}$. All three isotopes have six protons, but they have six, seven, and eight neutrons, respectively. Because natural carbon is a mixture of isotopes, the atomic mass we use for carbon is an \textit{average value} reflecting the average of the isotopes composing it.

The average atomic mass for carbon is computed as follows: It is known that natural carbon is composed of 98.89\% $^{12}\text{C}$ atoms and 1.11\% $^{13}\text{C}$ atoms. The amount of $^{14}\text{C}$ is negligibly small at this level of precision. Using the masses of $^{12}\text{C}$ (exactly 12\text{u}) and $^{13}\text{C}(13.003355\text{u})$, we can calculate the average atomic mass for natural carbon as follows:

\[
98.89\% \text{ of } 12\text{u} + 1.11\% \text{ of } 13.0034\text{u} = (0.9889)(12\text{u}) + (0.0111)(13.0034\text{u}) = 12.01\text{u}
\]

It is much easier to weigh out 600\text{hex} nuts than to count them one by one.

In this text we will call the average mass for an element the \textbf{average atomic mass} (the weighted average mass of the atoms in a naturally occurring element; also known as atomic weight) or, simply, the \textit{atomic mass} for that element.

Even though natural carbon does not contain a single atom with mass 12.01, for stoichiometric purposes, we can consider carbon to be composed of only one type of atom.
with a mass of 12.01. This enables us to count atoms of natural carbon by weighing a sample of carbon.

Recall from Section 3.1 that counting by weighing works if you know the average mass of the units being counted. Counting by weighing works just the same for atoms as for jelly beans. For natural carbon with an average mass of 12.01 atomic mass units, to obtain 1000 atoms would require weighing out 12,010 atomic mass units of natural carbon (a mixture of $^{12}\text{C}$ and $^{13}\text{C}$).

As in the case of carbon, the mass for each element listed in the table inside the front cover of the text is an average value based on the isotopic composition of the naturally occurring element. For instance, the mass listed for hydrogen (1.008) is the average mass for natural hydrogen, which is a mixture of $^1\text{H}$ and $^2\text{H}$ (deuterium). No atom of hydrogen actually has the mass 1.008.

In addition to being useful for determining accurate mass values for individual atoms, the mass spectrometer is used to determine the isotopic composition of a natural element. For example, when a sample of natural neon is injected into a mass spectrometer, the mass spectrum shown in Fig. 3.2 is obtained. The areas of the “peaks” or the heights of the bars indicate the relative abundances of $^{20}\text{Ne}$, $^{21}\text{Ne}$, and $^{22}\text{Ne}$ atoms.

**Figure 3.2**

(a) Neon gas glowing in a discharge tube. The relative intensities of the signals recorded when natural neon is injected into a mass spectrometer, represented in terms of (b) “peaks” and (c) a bar graph. The relative areas of the peaks are 0.9092$(^{20}\text{Ne})$, 0.00257$(^{21}\text{Ne})$, and 0.0882$(^{22}\text{Ne})$; natural neon is therefore 90.92%$^{20}\text{Ne}$, 0.257%$^{21}\text{Ne}$, and 8.82%$^{22}\text{Ne}$.

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**Example 3.1**

**The Average Mass of an Element**

When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 3.3 are obtained. Use these data to compute the average mass of natural copper. (The mass values for $^{63}\text{Cu}$ and $^{65}\text{Cu}$ are 62.93u and 64.93u, respectively.)
Figure 3.3
Mass spectrum of natural copper.

Copper nugget.

Solution

Where are we going?

To calculate the average mass of natural copper

What do we know?

- $^{63}\text{Cu}$ mass $= 62.93\text{u}$
- $^{65}\text{Cu}$ mass $= 64.93\text{u}$

How do we get there?

As shown by the graph, of every 100 atoms of natural copper, 69.09 are $^{63}\text{Cu}$ and 30.91 are $^{65}\text{Cu}$. Thus the mass of 100 atoms of natural copper is

\[
(69.09 \text{ atoms}) \left(2.93 \dfrac{\text{u}}{\text{atom}}\right) + (30.91 \text{ atoms}) \left(64.93 \dfrac{\text{u}}{\text{atom}}\right) = 6355\text{u}
\]
The average mass of a copper atom is
\[ \frac{6355u}{100\text{atoms}} = 63.55u/\text{atom} \]
This mass value is used in doing calculations involving the reactions of copper and is the value given in the table inside the front cover of this book.

**Reality Check**

When you finish a calculation, you should always check whether your answer makes sense. In this case our answer of 63.55u is between the masses of the atoms that make up natural copper. This makes sense. The answer could not be smaller than 62.93u or larger than 64.93u.

See Exercises 3.37 and 3.38.

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3.3 The Mole

AP Learning Objectives

- LO 1.4 The student is able to connect the number of particles, moles, mass, and volume of substances to one another, both qualitatively and quantitatively. [See SP 7.1; Essential knowledge 1.A.3]

Because samples of matter typically contain so many atoms, a unit of measure called the mole has been established for use in counting atoms. For our purposes, it is most convenient to define the mole (the number equal to the number of carbon atoms in exactly 12grams of pure $^{12}$C: Avogadro's number. One mole represents $6.022 \times 10^{23}$ units.) (abbreviated mol) as the number equal to the number of carbon atoms in exactly 12g of pure $^{12}$C. Techniques such as mass spectrometry, which count atoms very precisely, have been used to determine this number as $6.02214 \times 10^{23}$ (6.022 $\times 10^{23}$ will be sufficient for our purposes). This number is called Avogadro's number (the number of atoms in exactly 12grams of pure $^{12}$C, equal to $6.022 \times 10^{23}$) to honor his contributions to chemistry. One mole of something consists of $6.022 \times 10^{23}$ units of that substance. Just as a dozen eggs is 12 eggs, a mole of eggs is $6.022 \times 10^{23}$ eggs.

The magnitude of the number $6.022 \times 10^{23}$ is very difficult to imagine. To give you some
idea, 1 mole of seconds represents a span of time 4 million times as long as the earth has already existed, and 1 mole of marbles is enough to cover the entire earth to a depth of 50 miles! However, since atoms are so tiny, a mole of atoms or molecules is a perfectly manageable quantity to use in a reaction (Fig. 3.4).

**Figure 3.4**

One-mole samples of several elements.

Ken O’Donoghue © Cengage Learning

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**Critical Thinking**

What if you were offered $1 million to count from 1 to $6 \times 10^{23}$ at a rate of one number each second?

Determine your hourly wage. Would you do it? Could you do it?

---

How do we use the mole in chemical calculations? Recall that Avogadro's number is defined as the number of atoms in exactly 12 g of $^{12}$C. This means that 12 g of $^{12}$C contains $6.022 \times 10^{23}$ atoms. It also means that a 12.01-g sample of natural carbon contains $6.022 \times 10^{23}$ atoms (a mixture of $^{12}$C, $^{13}$C, and $^{14}$C atoms, with an average atomic mass of 12.01). Since the ratio of the masses of the samples (12 g/12.01 g) is the same as the ratio of the masses of the individual components (12 u/12.01 u), the two samples contain the same number of atoms ($6.022 \times 10^{23}$).
To be sure this point is clear, think of oranges with an average mass of 0.5 lb each and grapefruit with an average mass of 1.0 lb each. Any two sacks for which the sack of grapefruit weighs twice as much as the sack of oranges will contain the same number of pieces of fruit. The same idea extends to atoms. Compare natural carbon (average mass of 12.01 g) and natural helium (average mass of 4.003 g). A sample of 12.01 g of natural carbon contains the same number of atoms as 4.003 g of natural helium. Both samples contain 1 mole of atoms (6.022 × 10^23). Table 3.1 gives more examples that illustrate this basic idea.

<table>
<thead>
<tr>
<th>Element</th>
<th>Number of Atoms Present</th>
<th>Mass of Sample (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>6.022 × 10^{23}</td>
<td>26.98</td>
</tr>
<tr>
<td>Copper</td>
<td>6.022 × 10^{23}</td>
<td>63.55</td>
</tr>
<tr>
<td>Iron</td>
<td>6.022 × 10^{23}</td>
<td>55.85</td>
</tr>
<tr>
<td>Sulfur</td>
<td>6.022 × 10^{23}</td>
<td>32.07</td>
</tr>
<tr>
<td>Iodine</td>
<td>6.022 × 10^{23}</td>
<td>126.9</td>
</tr>
<tr>
<td>Mercury</td>
<td>6.022 × 10^{23}</td>
<td>200.6</td>
</tr>
</tbody>
</table>

Thus the mole is defined such that a sample of a natural element with a mass equal to the element’s atomic mass expressed in grams contains 1 mole of atoms. This definition also fixes the relationship between the atomic mass unit and the gram. Since 6.022 × 10^{23} atoms of carbon (each with a mass of 12 u) have a mass of 12 g, then

\[
(6.022 \times 10^{23} \text{ atoms}) \left(\frac{12 \text{ u}}{\text{atom}}\right) = 12 \text{ g}
\]

and

\[
6.022 \times 10^{23} \text{ u} = 1 \text{ g}
\]

This relationship can be used to derive the unit factor needed to convert between atomic mass units and grams.
Critical Thinking

What if you discovered Avogadro’s number was not $6.02 \times 10^{23}$ but $3.01 \times 10^{23}$? 

Would this affect the relative masses given on the periodic table? If so, how? If not, why not?

Interactive Example 3.2

**Determining the Mass of a Sample of Atoms**

Amercium is an element that does not occur naturally. It can be made in very small amounts in a device known as a *particle accelerator*. Compute the mass in grams of a sample of americium containing six atoms.

**Solution**

Where are we going?

To calculate the mass of six americium atoms

What do we know?

- Mass of 1 atom of Am = 243 u (from the periodic table inside the front cover)

How do we get there?

The mass of six atoms is

$$6 \text{ atoms} \times 243 \frac{u}{\text{atom}} = 1.46 \times 10^3 \text{u}$$

Using the relationship

$$6.022 \times 10^{23} \text{u} = 1 \text{g}$$

we write the conversion factor for converting atomic mass units to grams:

$$\frac{1 \text{g}}{6.022 \times 10^{23} \text{u}}$$

The mass of six americium atoms in grams is
Reality Check

Since this sample contains only six atoms, the mass should be very small as the amount $2.42 \times 10^{-21}$ g indicates.

See Exercise 3.45

To do chemical calculations, you must understand what the mole means and how to determine the number of moles in a given mass of a substance. These procedures are illustrated in Examples 3.3 and 3.4.

Interactive Example 3.3

Determining Moles of Atoms

Aluminum (Al) is a metal with a high strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

(left) Pure aluminum. (right) Aluminum alloys are used for many products used in our kitchens.

Solution

Where are we going?

To calculate the moles and number of atoms in a sample of Al

What do we know?
Sample contains 10.0g of Al

- Mass of 1 mole (6.022 x 10^23 atoms) of Al = 26.93g

**How do we get there?**

We can calculate the number of moles of Al in a 10.0-g sample as follows:

\[ \frac{10.0 \text{ g Al}}{26.98 \text{ g Al}} \times 1 \text{ mol Al} = 0.371 \text{ mol Al} \]

The number of atoms in 10.0g (0.371 mole) of aluminum is

\[ \frac{0.371 \text{ mol Al}}{1 \text{ mol Al}} \times 6.022 \times 10^{23} \text{ atoms} = 2.23 \times 10^{23} \text{ atoms} \]

**Reality Check**

One mole of Al has a mass of 26.98g and contains 6.022 x 10^23 atoms. Our sample is 10.0g, which is roughly 1/3 of 26.98. Thus the calculated amount should be on the order of 1/3 of 6 x 10^23, which it is.

See Exercise 3.46

---

**Interactive Example 3.4**

**Calculating Numbers of Atoms**

A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68mg. How many silicon (Si) atoms are present in the chip?

**Solution**

**Where are we going?**

To calculate the atoms of Si in the chip

**What do we know?**

- The chip has 5.68mg of Si
- Mass of 1 mole (6.022 x 10^23 atoms) of Si = 28.09g

**How do we get there?**

The strategy for doing this problem is to convert from milligrams of silicon to grams
of silicon, then to moles of silicon, and finally to atoms of silicon:

\[
\begin{align*}
5.68 \text{ mg Si} & \times \frac{1 \text{ g Si}}{1000 \text{ mg Si}} = 5.68 \times 10^{-3} \text{ g Si} \\
5.68 \times 10^{-3} \text{ g Si} & \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} = 2.02 \times 10^{-4} \text{ mol Si} \\
2.02 \times 10^{-4} \text{ mol Si} & \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol Si}} = 1.22 \times 10^{20} \text{ atoms}
\end{align*}
\]

**Reality Check**

Note that 5.68 mg of silicon is clearly much less than 1 mol of silicon (which has a mass of 28.09 g), so the final answer of 1.22 × 10²⁰ atoms (compared with 6.022 × 10²³ atoms) is in the right direction.

See Exercise 3.47

---

**Interactive Example 3.5**

**Calculating the Number of Moles and Mass**

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00 × 10²⁰ atoms and the mass of the sample.

Fragments of cobalt metal.

**Solution**

**Where are we going?**

To calculate the number of moles and the mass of a sample of Co

**What do we know?**
Sample contains \( 5.00 \times 10^{20} \) atoms of Co

**How do we get there?**

Note that the sample of \( 5.00 \times 10^{20} \) atoms of cobalt is less than \( 1 \text{ mole} \) \( (6.022 \times 10^{23} \text{ atoms}) \) of cobalt. What fraction of a mole it represents can be determined as follows:

\[
5.00 \times 10^{20} \text{ atoms Co} \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \text{ atoms Co}} = 8.30 \times 10^{-4} \text{ mol Co}
\]

Since the mass of \( 1 \text{ mole} \) of cobalt atoms is \( 58.93 \text{ g} \), the mass of \( 5.00 \times 10^{20} \) atoms can be determined as follows:

\[
8.30 \times 10^{-4} \text{ mol Co} \times \frac{58.93 \text{ g Co}}{1 \text{ mol Co}} = 4.89 \times 10^{-2} \text{ g Co}
\]

**Reality Check**

In this case the sample contains \( 5 \times 10^{20} \) atoms, which is approximately \( 1/1000 \) of a mole. Thus the sample should have a mass of about \( (1/1000)(58.93) \approx 0.06 \). Our answer of \( \sim 0.05 \) makes sense.

See Exercise 3.48

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**3.4 Molar Mass**

**AP Learning Objectives**

LO 1.4 The student is able to connect the number of particles, moles, mass, and volume of substances to one another, both qualitatively and quantitatively. [See SP 7.1; Essential knowledge 1.A.3]

A chemical compound is, ultimately, a collection of atoms. For example, methane (the major component of natural gas) consists of molecules that each contain one carbon and four hydrogen atoms (\( \text{CH}_4 \)). How can we calculate the mass of \( 1 \text{ mole} \) of methane; that is, what is the mass of \( 6.022 \times 10^{23} \text{ CH}_4 \) molecules? Since each \( \text{CH}_4 \) molecule contains one carbon atom and four hydrogen atoms, \( 1 \text{ mole} \) of \( \text{CH}_4 \) molecules contains \( 1 \text{ mole} \) of carbon atoms.
and 4 moles of hydrogen atoms. The mass of 1 mole of methane can be found by summing the masses of carbon and hydrogen present:

\[
\begin{align*}
\text{Mass of 1 mol C} & = 12.01 \text{ g} \\
\text{Mass of 4 mol H} & = 4 \times 1.008 \text{ g} \\
\text{Mass of 1 mol CH}_4 & = 16.04 \text{ g}
\end{align*}
\]

Because 16.04 g represents the mass of 1 mole of methane molecules, it makes sense to call it the molar mass for methane. Thus the molar mass (the mass in grams of one mole of molecules or formula units of a substance; also called molecular weight) of a substance is the mass in grams of 1 mole of the compound. Traditionally, the term molecular weight has been used for this quantity. However, we will use molar mass exclusively in this text. The molar mass of a known substance is obtained by summing the masses of the component atoms as we did for methane.

Methane is a molecular compound—its components are molecules. Many substances are ionic—they contain simple ions or polyatomic ions. Examples are \(\text{NaCl}\) (contains \(\text{Na}^+\) and \(\text{Cl}^-\)) and \(\text{CaCO}_3\) (contains \(\text{Ca}^{2+}\) and \(\text{CO}_3^{2-}\)). Because ionic compounds do not contain molecules, we need a special name for the fundamental unit of these materials. Instead of molecule, we use the term formula unit. Thus \(\text{CaCO}_3\) is the formula unit for calcium carbonate, and \(\text{NaCl}\) is the formula unit for sodium chloride.

### Interactive Example 3.6

**Calculating Molar Mass I**

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants. The formula for juglone is \(\text{C}_{10}\text{H}_6\text{O}_3\).

a. Calculate the molar mass of juglone.

b. A sample of \(1.56 \times 10^{-2}\) g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?
Solution

a. The molar mass is obtained by summing the masses of the component atoms. In 1 mole of juglone, there are 10 moles of carbon atoms, 6 moles of hydrogen atoms, and 3 moles of oxygen atoms:

\[
\begin{align*}
10 \text{ C} & : \quad 10 \times 12.01\text{g} = 120.1\text{g} \\
6 \text{ H} & : \quad 6 \times 1.008\text{g} = 6.048\text{g} \\
3 \text{ O} & : \quad 3 \times 16.00\text{g} = 48.00\text{g} \\
\text{Mass of } 1\text{mol C}_{10}\text{H}_6\text{O}_3 & = 174.1\text{g}
\end{align*}
\]

The mass of 1 mole of juglone is 174.1g, which is the molar mass.

b. The mass of 1 mole of this compound is 174.1g; thus 1.56 \times 10^{-2}\text{g} is much less than a mole. The exact fraction of a mole can be determined as follows:

\[
1.56 \times 10^{-2}\text{g juglone} \times \frac{1\text{ mol juglone}}{174.1\text{ g juglone}} = 8.96 \times 10^{-5}\text{mol juglone}
\]

See Exercises 3.51 through 3.54

Interactive Example 3.7

Calculating Molar Mass II

Calcium carbonate (\text{CaCO}_3), also called calcite, is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

a. Calculate the molar mass of calcium carbonate.

b. A certain sample of calcium carbonate contains 4.86 moles. What is
the mass in grams of this sample? What is the mass of the CO$_3^{2-}$ ions present?

A calcite (CaCO$_3$) crystal.

Charles D. Winters

**Solution**

a. Calcium carbonate is an ionic compound composed of Ca$^{2+}$ and CO$_3^{2-}$ ions. In 1 mole of calcium carbonate, there are 1 mole of Ca$^{2+}$ ions and 1 mole of CO$_3^{2-}$ ions. The molar mass is calculated by summing the masses of the components:

$$
\begin{align*}
1 \text{Ca}^{2+}: & \ 1 \times 40.08 \text{g} = 40.08 \text{g} \\
1 \text{CO}_3^{2-}: & \\
1 \text{C}: & \ 1 \times 12.01 \text{g} = 12.01 \text{g} \\
3 \text{O}: & \ 3 \times 16.00 \text{g} = 48.00 \text{g} \\
\text{Mass of} \ 1 \text{mol} \text{CaCO}_3 & = 100.09 \text{g}
\end{align*}
$$

Thus the mass of 1 mole of CaCO$_3$ (1 mole of Ca$^{2+}$ plus 1 mole of CO$_3^{2-}$) is 100.09 g. This is the molar mass.

b. The mass of 1 mole of CaCO$_3$ is 100.09 g. The sample contains nearly 5 moles, or close to 500 g. The exact amount is determined as follows:

$$
4.86 \text{ molCaCO}_3 \times \frac{100.09 \text{gCaCO}_3}{1 \text{ molCaCO}_3} = 486 \text{gCaCO}_3
$$

To find the mass of carbonate ions (CO$_3^{2-}$) present in this sample, we must realize that 4.86 moles of CaCO$_3$ contains 4.86 moles of Ca$^{2+}$ ions and 4.86 moles of CO$_3^{2-}$ ions. The mass of 1 mole of CO$_3^{2-}$ ions is

$$
\begin{align*}
1 \text{C}: & \ 1 \times 12.01 = 12.01 \text{g} \\
3 \text{O}: & \ 3 \times 16.00 = 48.00 \text{g} \\
\text{Mass of} \ 1 \text{mol} \text{CO}_3^{2-} & = 60.01 \text{g}
\end{align*}
$$
Thus the mass of \(4.86\text{ moles}\) of \(\text{CO}_3^{2-}\) ions is

\[
4.86 \text{ mol CO}_3^{2-} \times \frac{60.01 \text{ g CO}_3^{2-}}{1 \text{ mol CO}_3^{2-}} = 292 \text{ g CO}_3^{2-}
\]

See Exercises 3.55 through 3.58.

Interactive Example 3.8

**Molar Mass and Numbers of Molecules**

Isopentyl acetate \((\text{C}_7\text{H}_{14}\text{O}_2)\) is the compound responsible for the scent of bananas. A molecular model of isopentyl acetate is shown below. Interestingly, bees release about \(1\mu\text{g}(1 \times 10^{-6}\text{g})\) of this compound when they sting. The resulting scent attracts other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting? How many atoms of carbon are present?

Isopentyl acetate is released when a bee stings.
Solution

Where are we going?

To calculate the number of molecules of isopentyl acetate and the number of carbon atoms in a bee sting

What do we know?

- Mass of isopentyl acetate in a typical bee sting is 1 microgram \(= 1 \times 10^{-6} \text{ g} \)

How do we get there?

Since we are given a mass of isopentyl acetate and want to find the number of molecules, we must first compute the molar mass of \( \text{C}_7\text{H}_{14}\text{O}_2 \):

\[
\begin{align*}
7 \text{ mol} & \times 12.01 \frac{\text{g}}{\text{mol}} = 84.07 \text{ g C} \\
14 \text{ mol} & \times 1.008 \frac{\text{g}}{\text{mol}} = 14.11 \text{ g H} \\
2 \text{ mol} & \times 16.00 \frac{\text{g}}{\text{mol}} = 32.00 \frac{\text{g O}}{130.18 \text{ g}}
\end{align*}
\]

This means that 1 mole of isopentyl acetate (\(6.022 \times 10^{23}\) molecules) has a mass of 130.18 g.

To find the number of molecules released in a sting, we must first determine the number of moles of isopentyl acetate in \(1 \times 10^{-6} \text{ g}\):

\[
1 \times 10^{-6} \text{ gC}_7\text{H}_{14}\text{O}_2 \times \frac{1 \text{ molC}_7\text{H}_{14}\text{O}_2}{130.18 \text{ gC}_7\text{H}_{14}\text{O}_2} = 8 \times 10^{-9} \text{ molC}_7\text{H}_{14}\text{O}_2
\]

Since 1 mole is \(6.022 \times 10^{23}\) units, we can determine the number of molecules:

\[
8 \times 10^{-9} \text{ molC}_7\text{H}_{14}\text{O}_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ molC}_7\text{H}_{14}\text{O}_2} = 5 \times 10^{15} \text{ molecules}
\]

To determine the number of carbon atoms present, we must multiply the number
One of the great rewards of studying chemistry is to become a good problem solver. Being able to solve complex problems is a talent that will serve you well in all walks of life. It is our purpose in this text to help you learn to solve problems in a flexible, creative way based on understanding the fundamental ideas of chemistry. We call this approach **conceptual problem solving**.

The ultimate goal is to be able to solve new problems (that is, problems you have not seen before) on your own. In this text we will provide problems and offer solutions by explaining how to think about the problems. While the answers to these problems are important, it is perhaps even more important to understand the process—the thinking necessary to get the answer. Although at first we will be solving the problem for you, do not take a passive role. While studying the solution, it is crucial that you interactively think through the problem with us. Do not skip the discussion and jump to the answer. Usually, the solution will involve asking a series of questions. Make sure that you understand each step in the process. This active approach should apply to problems outside of chemistry as well. For example, imagine riding with someone in a car to an unfamiliar destination. If your goal is simply to have the other person get you to that destination, you will probably not pay much attention to how to get there (passive), and if you have to find this same place
in the future on your own, you probably will not be able to do it. If, however, your goal is to learn how to get there, you would pay attention to distances, signs, and turns (active). This is how you should read the solutions in the text (and the text in general).

While actively studying our solutions to problems is helpful, at some point you will need to know how to think through these problems on your own. If we help you too much as you solve a problem, you won’t really learn effectively. If we always “drive,” you won’t interact as meaningfully with the material. Eventually you need to learn to drive yourself. We will provide more help at the beginning of the text and less as we proceed to later chapters.

Pigeonholes can be used for sorting and classifying objects like mail.

There are two fundamentally different ways you might use to approach a problem. One way emphasizes memorization. We might call this the “pigeonholing method.” In this approach, the first step is to label the problem—to decide in which pigeonhole it fits. The pigeonholing method requires that we provide you with a set of steps that you memorize and store in the appropriate slot for each different problem you encounter. The difficulty with this method is that it requires a new pigeonhole each time a problem is changed by even a small amount.

Consider the driving analogy again. Suppose you have memorized how to drive from your house to the grocery store. Do you know how to drive back from the grocery store to your house? Not necessarily. If you have only memorized the directions and do not understand fundamental principles such as “I traveled north to get to the store, so my house is south of the store,” you may find yourself stranded. In a more complicated example, suppose you know how to get from your house to the store (and back) and from your house to the library (and back). Can you get from the library to the store without having to go back home? Probably not if you have only memorized directions and you do not have a “big picture” of where your house, the store, and the library are relative to one another.

The second approach is conceptual problem solving, in which we help you get the “big picture”—a real understanding of the situation. This approach to problem solving looks within the problem for a solution. In this method we assume that the problem is a new one, and we let the problem guide us as we solve it. In this approach we ask a series of questions as we proceed and use our knowledge of fundamental principles to answer these questions. Learning this approach requires some patience, but the reward for learning to solve
problems this way is that we become an effective solver of any new problem that confronts us in daily life or in our work in any field. In summary, instead of looking outside the problem for a memorized solution, we will look inside the problem and let the problem help us as we proceed to a solution.

As we have seen in problems we have already considered, there are several organizing principles to help you become a creative problem solver. Although we have been using these ideas in earlier problems, let’s review and expand on them. Because as we progress in our study of chemistry the problems become more complicated, we will need to rely on this approach even more.

1. We need to read the problem and decide on the final goal. Then we sort through the facts given, focusing on the key words and often drawing a diagram of the problem. In this part of the analysis we need to state the problem as simply and as visually as possible. We could summarize this entire process as “Where are we going?”

2. In order to reach our final goal, we need to decide where to start. For example, in a stoichiometry problem we always start with the chemical reaction. Then we ask a series of questions as we proceed, such as, “What are the reactants and products?” “What is the balanced equation?” “What are the amounts of the reactants?” and so on. Our understanding of the fundamental principles of chemistry will enable us to answer each of these simple questions and eventually will lead us to the final solution. We might summarize this process as “How do we get there?”

3. Once we get the solution of the problem, then we ask ourselves, “Does it make sense?” That is, does our answer seem reasonable? We call this the Reality Check. It always pays to check your answer.

Using a conceptual approach to problem solving will enable you to develop real confidence as a problem solver. You will no longer panic when you see a problem that is different in some ways from those you have solved in the past. Although you might be frustrated at times as you learn this method, we guarantee that it will pay dividends later and should make your experience with chemistry a positive one that will prepare you for any career you choose.

To summarize, one of our major goals in this text is to help you become a creative problem solver. We will do this by, at first, giving you lots of guidance in how to solve problems. We will “drive,” but we hope you will be paying attention instead of just “riding along.” As we move forward, we will gradually shift more of the responsibility to you. As you gain confidence in letting the problem guide you, you will be amazed at how effective you can be at solving some really complex problems — just like the ones you will confront in “real life.”
There are two common ways of describing the composition of a compound: in terms of the numbers of its constituent atoms and in terms of the percentages (by mass) of its elements. We can obtain the mass percents of the elements from the formula of the compound by comparing the mass of each element present in 1 mole of the compound to the total mass of 1 mole of the compound.

For example, for ethanol, which has the formula \( \text{C}_2\text{H}_5\text{OH} \), the mass of each element present and the molar mass are obtained as follows:

Mass of C \( = 2 \text{ mol} \times 12.01 \frac{\text{g}}{\text{mol}} = 24.02\text{g} \)

Mass of H \( = 6 \text{ mol} \times 1.008 \frac{\text{g}}{\text{mol}} = 6.048\text{g} \)

Mass of O \( = 1 \text{ mol} \times 16.00 \frac{\text{g}}{\text{mol}} = 16.00\text{g} \)

Mass of 1 mol \( \text{C}_2\text{H}_5\text{OH} \) \( = 46.07\text{g} \)

The **mass percent** (the percent by mass of a component of a mixture or of a given element in a compound) (often called the **weight percent**) of carbon in ethanol can be computed by comparing the mass of carbon in 1 mole of ethanol to the total mass of 1 mole of ethanol and multiplying the result by 100:

\[
\text{Mass percent of C} = \left( \frac{\text{mass of C in 1 mol \text{C}_2\text{H}_5\text{OH}}}{\text{mass of 1 mol \text{C}_2\text{H}_5\text{OH}}} \right) \times 100\% \\
= \left( \frac{24.02\text{g}}{46.07\text{g}} \right) \times 100\% = 52.14\%
\]

The mass percents of hydrogen and oxygen in ethanol are obtained in a similar manner:
Interactive Example 3.9

Calculating Mass Percent

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula \((C_{10}H_{14}O)\) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

Solution

Where are we going?

To find the mass percent of each element in carvone

What do we know?

- Molecular formula is \(C_{10}H_{14}O\)

What information do we need to find the mass percent?

- Mass of each element (we’ll use 1 mole of carvone)
- Molar mass of carvone

Mass percent of H

\[
\text{Mass percent of H} = \frac{\text{mass of H in 1 mol C}_2\text{H}_5\text{OH}}{\text{mass of 1 mol C}_2\text{H}_5\text{OH}} \times 100\% \\
= \frac{6.048 \text{ g}}{46.07 \text{ g}} \times 100\% = 13.13\%
\]

Mass percent of O

\[
\text{Mass percent of O} = \frac{\text{mass of O in 1 mol C}_2\text{H}_5\text{OH}}{\text{mass of 1 mol C}_2\text{H}_5\text{OH}} \times 100\% \\
= \frac{16.00 \text{ g}}{46.07 \text{ g}} \times 100\% = 34.73\%
\]

Reality Check Notice that the percentages add up to 100.00%; this provides a check that the calculations are correct.
How do we get there?

What is the mass of each element in 1 mole of $C_{10}H_{14}O$?

$$\text{Mass of C in 1 mol} = 10 \text{ mol} \times 12.01 \text{ g mol}^{-1} = 120.1 \text{ g}$$

$$\text{Mass of H in 1 mol} = 14 \text{ mol} \times 1.008 \text{ g mol}^{-1} = 14.11 \text{ g}$$

$$\text{Mass of O in 1 mol} = 1 \text{ mol} \times 16.00 \text{ g mol}^{-1} = 16.00 \text{ g}$$

What is the molar mass of $C_{10}H_{14}O$?

$$120.1 \text{ g C} + 14.11 \text{ g H} + 16.00 \text{ g O} = 150.2 \text{ g}$$

$C_{10}$ + $H_{14}$ + O = $C_{10}H_{14}O$

What is the mass percent of each element?

We find the fraction of the total mass contributed by each element and convert it to a percentage:

- Mass percent of C = \( \frac{120.1 \text{ g C}}{150.2 \text{ g} C_{10}H_{14}O} \times 100\% = 79.96\% \)
- Mass percent of H = \( \frac{14.11 \text{ g H}}{150.2 \text{ g} C_{10}H_{14}O} \times 100\% = 9.394\% \)
- Mass percent of O = \( \frac{16.00 \text{ g O}}{150.2 \text{ g} C_{10}H_{14}O} \times 100\% = 10.65\% \)

Reality Check

Sum the individual mass percent values—they should total to 100% within round-off errors. In this case, the percentages add up to 100.00%.

See Exercises 3.73 and 3.74
3.7 Determining the Formula of a Compound

When a new compound is prepared, one of the first items of interest is the formula of the compound. This is most often determined by taking a weighed sample of the compound and either decomposing it into its component elements or reacting it with oxygen to produce substances such as $\text{CO}_2$, $\text{H}_2\text{O}$, and $\text{N}_2$, which are then collected and weighed. A device for doing this type of analysis is shown in Fig. 3.5. The results of such analyses provide the mass of each type of element in the compound, which can be used to determine the mass percent of each element.

Figure 3.5

A schematic diagram of the combustion device used to analyze substances for carbon and hydrogen. The sample is burned in the presence of excess oxygen, which converts all its carbon to carbon dioxide and all its hydrogen to water. These products are collected by absorption using appropriate materials, and their amounts are determined by measuring the increase in masses of the absorbents.

We will see how information of this type can be used to compute the formula of a compound. Suppose a substance has been prepared that is composed of carbon, hydrogen,
and nitrogen. When 0.1156 g of this compound is reacted with oxygen, 0.1638 g of carbon dioxide (CO₂) and 0.1676 g of water (H₂O) are collected. Assuming that all the carbon in the compound is converted to CO₂, we can determine the mass of carbon originally present in the 0.1156 g sample. To do this, we must use the fraction (by mass) of carbon in CO₂. The molar mass of CO₂ is

\[
\text{C: } 1 \text{ mol} \times 12.01 \frac{\text{g}}{\text{mol}} = 12.01 \text{ g}
\]

\[
\text{O: } 2 \text{ mol} \times 16.00 \frac{\text{g}}{\text{mol}} = 32.00 \text{ g}
\]

\[
\text{Molar Mass of CO}_2 = 44.01 \text{ g/mol}
\]

The fraction of carbon present by mass is

\[
\frac{\text{Mass of C}}{\text{Total mass of CO}_2} = \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2}
\]

This factor can now be used to determine the mass of carbon in 0.1638 g of CO₂:

\[
0.1638 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.04470 \text{ g C}
\]

Remember that this carbon originally came from the 0.1156 g sample of unknown compound. Thus the mass percent of carbon in this compound is

\[
\frac{0.04470 \text{ g C}}{0.1156 \text{ g compound}} \times 100\% = 38.67\% \text{ C}
\]

The same procedure can be used to find the mass percent of hydrogen in the unknown compound. We assume that all the hydrogen present in the original 0.1156 g of compound was converted to H₂O. The molar mass of H₂O is 18.02 g, and the fraction of hydrogen by mass in H₂O is

\[
\frac{\text{Mass of H}}{\text{Mass of H}_2\text{O}} = \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}}
\]

Therefore, the mass of hydrogen in 0.1676 g of H₂O is
The mass percent of hydrogen in the compound is

\[
\frac{0.01875 \text{ g H}}{0.1156 \text{ g compound}} \times 100\% = 16.22\% \text{ H}
\]

The unknown compound contains only carbon, hydrogen, and nitrogen. So far we have determined that it is 38.67\% carbon and 16.22\% hydrogen. The remainder must be nitrogen:

\[
100.00\% - (38.67\% + 16.22\%) = 45.11\% \text{ N}
\]

We have determined that the compound contains 38.67\% carbon, 16.22\% hydrogen, and 45.11\% nitrogen. Next we use these data to obtain the formula.

Since the formula of a compound indicates the numbers of atoms in the compound, we must convert the masses of the elements to numbers of atoms. The easiest way to do this is to work with 100.00 g of the compound. In the present case, 38.67\% carbon by mass means 38.67 g of carbon per 100.00 g of compound; 16.22\% hydrogen means 16.22 g of hydrogen per 100.00 g of compound; and so on. To determine the formula, we must calculate the number of carbon atoms in 38.67 g of carbon, the number of hydrogen atoms in 16.22 g of hydrogen, and the number of nitrogen atoms in 45.11 g of nitrogen. We can do this as follows:

\[
\begin{align*}
38.67 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} & = 3.220 \text{ mol C} \\
16.22 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} & = 16.09 \text{ mol H} \\
45.11 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} & = 3.220 \text{ mol N}
\end{align*}
\]

Thus 100.00 g of this compound contains 3.220 moles of carbon atoms, 16.09 moles of hydrogen atoms, and 3.220 moles of nitrogen atoms.

We can find the smallest whole-number ratio of atoms in this compound by dividing each of the mole values above by the smallest of the three:

\[
\begin{align*}
\text{C:} & \quad \frac{3.220}{3.220} = 1.000 = 1 \\
\text{H:} & \quad \frac{16.09}{3.220} = 4.997 = 5 \\
\text{N:} & \quad \frac{3.220}{3.220} = 1.000 = 1
\end{align*}
\]

Thus the formula might well be CH\(_5\)N. However, it also could be C\(_2\)H\(_{10}\)N\(_2\) or C\(_3\)H\(_{15}\)N\(_3\), and so on—that is, some multiple of the smallest whole-number ratio. Each of these alternatives also has the correct relative numbers of atoms. That is, any molecule that can be represented as (CH\(_5\)N)\(_n\), where \(n\) is an integer, has the empirical formula (the...
simplest whole-number ratio of atoms in a compound.) CH₅N. To be able to specify the exact formula of the molecule involved, the molecular formula (the exact formula of a molecule, giving the types of atoms and the number of each type.), we must know the molar mass. 

Suppose we know that this compound with empirical formula CH₅N has a molar mass of 31.06 g/mol. How do we determine which of the possible choices represents the molecular formula? Since the molecular formula is always a whole-number multiple of the empirical formula, we must first find the empirical formula mass for CH₅N:

\[
\begin{align*}
1 \text{ C} & : 1 \times 12.01 \text{g} = 12.01 \text{g} \\
5 \text{ H} & : 5 \times 1.008 \text{g} = 5.040 \text{g} \\
1 \text{ N} & : 1 \times 14.01 \text{g} = 14.01 \text{g} \\
\text{Formula mass of CH₅N} & = 31.06 \text{ g/mol}
\end{align*}
\]

This is the same as the known molar mass of the compound. Thus in this case the empirical formula and the molecular formula are the same; this substance consists of molecules with the formula CH₅N. It is quite common for the empirical and molecular formulas to be different; some examples where this is the case are shown in Fig. 3.6.

**Figure 3.6**

Examples of substances whose empirical and molecular formulas differ. Notice that molecular formula = (empirical formula)_n, where n is an integer.

C₆H₆ = (CH)₆  \quad S₈ = (S)₈  \quad C₆H₁₂O₆ = (CH₂O)₆

**Problem-Solving Strategy**

**Empirical Formula Determination**

- Since mass percentage gives the number of grams of a particular element per 100 g of compound, base the calculation on 100 g of compound. Each percent will then represent the mass in grams of that element.

- Determine the number of moles of each element present in 100 g of compound using the atomic masses of the elements present.

- Divide each value of the number of moles by the smallest of the values. If
each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.

- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

Critical Thinking

One part of the problem-solving strategy for empirical formula determination is to base the calculation on 100g of compound. What if you chose a mass other than 100g? Would this work? What if you chose to base the calculation on 100 moles of compound? Would this work?

Problem-Solving Strategy

**Determining Molecular Formula from Empirical Formula**

- Obtain the empirical formula.

- Compute the mass corresponding to the empirical formula.

- Calculate the ratio:

\[
\frac{\text{Molar mass}}{\text{Empirical formula mass}}
\]

- The integer from the previous step represents the number of empirical formula units in one molecule. When the empirical formula subscripts are multiplied by this integer, the molecular formula results. This procedure is summarized by the equation:

\[
\text{Molecular formula} = \text{empirical formula} \times \frac{\text{molar mass}}{\text{empirical formula mass}}
\]

Interactive Example 3.10

**Determining Empirical and Molecular Formulas I**
Determine the empirical and molecular formulas for a compound that gives the following percentages on analysis (in mass percents):

\[ 71.65\% \text{Cl} \quad 24.27\% \text{C} \quad 4.07\% \text{H} \]

The molar mass is known to be \( 98.96\text{g/mol} \).

**Solution**

**Where are we going?**

To find the empirical and molecular formulas for the given compound.

**What do we know?**

- Percent of each element
- Molar mass of the compound is \( 98.96\text{g/mol} \)

**What information do we need to find the empirical formula?**

- Mass of each element in \( 100.00\text{g} \) of compound
- Moles of each element

**How do we get there?**

**What is the mass of each element in \( 100.00\text{g} \) of compound?**

\[
\text{Cl} \quad 71.65\text{g} \quad \text{C} \quad 24.27\text{g} \quad \text{H} \quad 4.07\text{g}
\]

**What are the moles of each element in \( 100.00\text{g} \) of compound?**

\[
\begin{align*}
71.65\ \text{g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} &= 2.021 \text{ mol Cl} \\
24.27\ \text{g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 2.021 \text{ mol C} \\
4.07\ \text{g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} &= 4.04 \text{ mol H}
\end{align*}
\]

**What is the empirical formula for the compound?**

Dividing each mole value by 2.021 (the smallest number of moles present), we find the empirical formula \( \text{ClC}_2 \).

**What is the molecular formula for the compound?**

Compare the empirical formula mass to the molar mass.
This substance is composed of molecules with the formula \( \text{Cl}_2\text{C}_2\text{H}_4 \).

Note: The method we use here allows us to determine the molecular formula of a compound but not its structural formula. The compound \( \text{Cl}_2\text{C}_2\text{H}_4 \) is called dichloroethane. There are two forms of this compound, shown in Fig. 3.7. The form at the bottom was formerly used as an additive in leaded gasoline.

**Figure 3.7**
The two forms of dichloroethane.

See Exercise 3.87 and 3.88

---

**Interactive Example 3.11**

**Determining Empirical and Molecular Formulas II**

A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a molar mass of 283.88g/mol. What are the compound’s empirical and molecular formulas?

**Solution**

**Where are we going?**

To find the empirical and molecular formulas for the given compound

**What do we know?**
Percent of each element

Molar mass of the compound is 283.88g/mol

What information do we need to find the empirical formula?

- Mass of each element in 100.00g of compound
- Moles of each element

**How do we get there?**

What is the mass of each element in 100.00g of compound?

P 43.64 g  O 56.36g

What are the moles of each element in 100.00g of compound?

\[
\begin{align*}
43.64 \text{ g P} & \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 1.409 \text{ mol P} \\
56.36 \text{ g O} & \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 3.523 \text{ mol O}
\end{align*}
\]

What is the empirical formula for the compound?

Dividing each mole value by the smaller one gives

\[
\frac{1.409}{1.409} = 1\text{P} \quad \text{and} \quad \frac{3.523}{1.409} = 2.5\text{O}
\]

This yields the formula PO\(_{2.5}\). Since compounds must contain whole numbers of atoms, the empirical formula should contain only whole numbers. To obtain the simplest set of whole numbers, we multiply both numbers by 2 to give the empirical formula P\(_2\)O\(_5\).

What is the molecular formula for the compound?

Compare the empirical formula mass to the molar mass.

Empirical formula mass = 141.94g/mol (Confirm this!)

Molar mass is given = 283.88g/mol

\[
\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{283.88}{141.94} = 2
\]

- The molecular formula is (P\(_2\)O\(_5\))\(_2\), or P\(_4\)O\(_{10}\).

Note: The structural formula for this interesting compound is given in Fig. 3.8.

**Figure 3.8**
The structure of $P_4O_{10}$ Note that some of the oxygen atoms act as “bridges” between the phosphorus atoms. This compound has a great affinity for water and is often used as a desiccant, or drying agent.

See Exercise 3.89

In Examples 3.10 and 3.11 we found the molecular formula by comparing the empirical formula mass with the molar mass. There is an alternate way to obtain the molecular formula. For example, in Example 3.10 we know the molar mass of the compound is $98.96\,\text{g/mol}$. This means that 1 mole of the compound weighs 98.96 g. Since we also know the mass percentages of each element, we can compute the mass of each element present in 1 mole of compound:

\[
\begin{align*}
\text{Chlorine:} & \quad \frac{71.65\,\text{g Cl}}{100.0\,\text{g compound}} \times \frac{98.96\,\text{g}}{\text{mol}} = \frac{70.90\,\text{g Cl}}{\text{mol compound}} \\
\text{Carbon:} & \quad \frac{24.27\,\text{g C}}{100.0\,\text{g compound}} \times \frac{98.96\,\text{g}}{\text{mol}} = \frac{24.02\,\text{g C}}{\text{mol compound}} \\
\text{Hydrogen:} & \quad \frac{4.07\,\text{g H}}{100.0\,\text{g compound}} \times \frac{98.96\,\text{g}}{\text{mol}} = \frac{4.03\,\text{g H}}{\text{mol compound}}
\end{align*}
\]

Now we can compute moles of atoms present per mole of compound:

\[
\begin{align*}
\text{Chlorine:} & \quad \frac{70.90\,\text{g Cl}}{\text{mol compound}} \times \frac{1\text{mol Cl}}{35.45\,\text{g Cl}} = \frac{2.000\text{mol Cl}}{\text{mol compound}} \\
\text{Carbon:} & \quad \frac{24.02\,\text{g C}}{\text{mol compound}} \times \frac{1\text{mol C}}{12.01\,\text{g C}} = \frac{2.000\text{mol C}}{\text{mol compound}} \\
\text{Hydrogen:} & \quad \frac{4.03\,\text{g H}}{\text{mol compound}} \times \frac{1\text{mol H}}{1.008\,\text{g H}} = \frac{4.00\text{mol H}}{\text{mol compound}}
\end{align*}
\]

Thus 1 mole of the compound contains 2 moles of Cl atoms, 2 moles of C atoms, and 4 moles of H atoms, and the molecular formula is $\text{Cl}_2\text{C}_2\text{H}_4$, as obtained in Example 3.10.

Problem-Solving Strategy
Determining Molecular Formula from Mass Percent and Molar Mass

- Using the mass percentages and the molar mass, determine the mass of each element present in 1 mole of compound.

- Determine the number of moles of each element present in 1 mole of compound.

- The integers from the previous step represent the subscripts in the molecular formula.

Interactive Example 3.12

Determining a Molecular Formula

Coffee, a stimulant found in coffee, tea, and chocolate, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a molar mass of 194.2 g/mol. Determine the molecular formula of caffeine.

Solution

Where are we going?

To find the molecular formula for caffeine

What do we know?

- Percent of each element
Molar mass of caffeine is 194.2g/mol

What information do we need to find the molecular formula?

- Mass of each element (in 1 mole of caffeine)
- Mole of each element (in 1 mole of caffeine)

How do we get there?

What is the mass of each element in 1 mole (194.2g) of caffeine?

\[
\begin{align*}
\frac{49.48 \text{ g C}}{100.0 \text{ g caffeine}} \times \frac{194.2 \text{ g}}{\text{mol}} &= \frac{96.09 \text{ g C}}{\text{mol caffeine}} \\
\frac{5.15 \text{ g H}}{100.0 \text{ g caffeine}} \times \frac{194.2 \text{ g}}{\text{mol}} &= \frac{10.0 \text{ g H}}{\text{mol caffeine}} \\
\frac{28.87 \text{ g N}}{100.0 \text{ g caffeine}} \times \frac{194.2 \text{ g}}{\text{mol}} &= \frac{56.07 \text{ g N}}{\text{mol caffeine}} \\
\frac{16.49 \text{ g O}}{100.0 \text{ g caffeine}} \times \frac{194.2 \text{ g}}{\text{mol}} &= \frac{32.02 \text{ g O}}{\text{mol caffeine}}
\end{align*}
\]

What are the moles of each element in 1 mole of caffeine?

\[
\begin{align*}
\text{Carbon:} & \quad \frac{96.09 \text{ g C}}{\text{mol caffeine}} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = \frac{8.001 \text{ mol C}}{\text{mol caffeine}} \\
\text{Hydrogen:} & \quad \frac{10.0 \text{ g H}}{\text{mol caffeine}} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = \frac{9.92 \text{ mol H}}{\text{mol caffeine}} \\
\text{Nitrogen:} & \quad \frac{56.07 \text{ g N}}{\text{mol caffeine}} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = \frac{4.002 \text{ mol N}}{\text{mol caffeine}} \\
\text{Oxygen:} & \quad \frac{32.02 \text{ g O}}{\text{mol caffeine}} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = \frac{2.001 \text{ mol O}}{\text{mol caffeine}}
\end{align*}
\]

Rounding the numbers to integers gives the molecular formula for caffeine: \( \text{C}_8\text{H}_{10}\text{N}_4\text{O}_2 \).

See Exercise 3.90
3.8 Chemical Equations

AP Learning Objectives

LO 1.18 The student is able to apply conservation of atoms to the rearrangement of atoms in various processes. [See SP 1.4; Essential knowledge 1.E.2]

LO 3.1 Students can translate among macroscopic observations of change, chemical equations, and particle views. [See SP 1.5, 7.1; Essential knowledge components of 3.A–3.C]

Chemical Reactions

A chemical change involves a reorganization of the atoms in one or more substances. For example, when the methane (\( \text{CH}_4 \)) in natural gas combines with oxygen (\( \text{O}_2 \)) in the air and burns, carbon dioxide (\( \text{CO}_2 \)) and water (\( \text{H}_2\text{O} \)) are formed. This process is represented by a chemical equation (a representation of a chemical reaction showing the relative numbers of reactant and product molecules) with the reactants (a starting substance in a chemical reaction. It appears to the left of the arrow in a chemical equation.) (here methane and oxygen) on the left side of an arrow and the products (a substance resulting from a chemical reaction. It is shown to the right of the arrow in a chemical equation.) (carbon dioxide and water) on the right side:

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Notice that the atoms have been reorganized. Bonds have been broken, and new ones have been formed. It is important to recognize that in a chemical reaction, atoms are neither created nor destroyed. All atoms present in the reactants must be accounted for among the products. In other words, there must be the same number of each type of atom on the products side and on the reactants side of the arrow. Making sure that this rule is obeyed is called balancing a chemical equation (making sure that all atoms present in the reactants
are accounted for among the products) for a reaction.

The equation (shown above) for the reaction between \( \text{CH}_4 \) and \( \text{O}_2 \) is not balanced. We can see this from the following representation of the reaction:

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Notice that the number of oxygen atoms (in \( \text{O}_2 \)) on the left of the arrow is two, while on the right there are three \( \text{O} \) atoms (in \( \text{CO}_2 \) and \( \text{H}_2\text{O} \)). Also, there are four hydrogen atoms (in \( \text{CH}_4 \)) on the left and only two (in \( \text{H}_2\text{O} \)) on the right. Remember that a chemical reaction is simply a rearrangement of the atoms (a change in the way they are organized). Atoms are neither created nor destroyed in a chemical reaction. Thus the reactants and products must occur in numbers that give the same number of each type of atom among both the reactants and products. Simple trial and error will allow us to figure this out for the reaction of methane with oxygen. The needed numbers of molecules are

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]

Notice that now we have the same number of each type of atom represented among the reactants and the products.

We can represent the preceding situation in a shorthand manner by the following chemical equation:

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}
\]

We can check that the equation is balanced by comparing the number of each type of atom on both sides:

\[
\begin{array}{ccc}
\text{CH}_4 & + & 2\text{O}_2 \\
\uparrow & & \uparrow \\
1 \text{C} & 4 \text{H} & 4 \text{O} \\
\downarrow & & \downarrow \\
\text{CO}_2 & + & 2\text{H}_2\text{O} \\
\uparrow & & \uparrow \\
1 \text{C} & 4 \text{H} & 4 \text{O} \\
\end{array}
\]

To summarize, we have

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>1( \text{C} )</td>
<td>1( \text{C} )</td>
</tr>
<tr>
<td>4( \text{H} )</td>
<td>4( \text{H} )</td>
</tr>
<tr>
<td>4( \text{O} )</td>
<td>4( \text{O} )</td>
</tr>
</tbody>
</table>
Chapter 3: Stoichiometry The Meaning of a Chemical Equation

The chemical equation for a reaction gives two important types of information: the nature of the reactants and products and the relative numbers of each.

The reactants and products in a specific reaction must be identified by experiment. Besides specifying the compounds involved in the reaction, the equation often gives the physical states of the reactants and products:

<table>
<thead>
<tr>
<th>State</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid</td>
<td>(s)</td>
</tr>
<tr>
<td>Liquid</td>
<td>(l)</td>
</tr>
<tr>
<td>Gas</td>
<td>(g)</td>
</tr>
<tr>
<td>Dissolved in water (in aqueous solution)</td>
<td>(aq)</td>
</tr>
</tbody>
</table>

For example, when hydrochloric acid in aqueous solution is added to solid sodium hydrogen carbonate, the products carbon dioxide gas, liquid water, and sodium chloride (which dissolves in the water) are formed:

\[ \text{HCl}(aq) + \text{NaHCO}_3(s) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

The relative numbers of reactants and products in a reaction are indicated by the coefficients in the balanced equation. (The coefficients can be determined because we know that the same number of each type of atom must occur on both sides of the equation.) For example, the balanced equation

\[ \text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g) \]

can be interpreted in several equivalent ways, as shown in Table 3.2. Note that the total mass is 80g for both reactants and products. We expect the mass to remain constant, since chemical reactions involve only a rearrangement of atoms. Atoms, and therefore mass, are conserved in a chemical reaction.
Table 3.2

Information Conveyed by the Balanced Equation for the Combustion of Methane

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₄(g) + 2O₂(g)</td>
<td>CO₂(g) + 2H₂O(g)</td>
</tr>
<tr>
<td>1molecule + 2molecules</td>
<td>1molecule + 2molecules</td>
</tr>
<tr>
<td>1mole + 2moles</td>
<td>1mole + 2moles</td>
</tr>
<tr>
<td>6.022 × 10²³ molecules + 2(6.022 × 10²³ molecules)</td>
<td>6.022 × 10²³ molecules + 2(6.022 × 10²³ molecules)</td>
</tr>
<tr>
<td>16g + 2(32g)</td>
<td>44g + 2(18g)</td>
</tr>
<tr>
<td>80g reactants</td>
<td>80g products</td>
</tr>
</tbody>
</table>

From this discussion you can see that a balanced chemical equation gives you a great deal of information.

Chapter 3: Stoichiometry: 3.9 Balancing Chemical Equations
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3.9 Balancing Chemical Equations

AP Learning Objectives

LO 1.17  The student is able to express the law of conservation of mass quantitatively and qualitatively using symbolic representations and particulate drawings. [See SP 1.5; Essential knowledge 1.E.1]

LO 1.18  The student is able to apply conservation of atoms to the rearrangement of atoms in various processes. [See SP 1.4; Essential knowledge 1.E.2]
An unbalanced chemical equation is of limited use. Whenever you see an equation, you should ask yourself whether it is balanced. The principle that lies at the heart of the balancing process is that atoms are conserved in a chemical reaction. The same number of each type of atom must be found among the reactants and products. It is also important to recognize that the identities of the reactants and products of a reaction are determined by experimental observation. For example, when liquid ethanol is burned in the presence of sufficient oxygen gas, the products are always carbon dioxide and water. When the equation for this reaction is balanced, the identities of the reactants and products must not be changed. The formulas of the compounds must never be changed in balancing a chemical equation. That is, the subscripts in a formula cannot be changed, nor can atoms be added or subtracted from a formula.

Critical Thinking

What if a friend was balancing chemical equations by changing the values of the subscripts instead of using the coefficients? How would you explain to your friend that this was the wrong thing to do?

Most chemical equations can be balanced by inspection, that is, by trial and error. It is always best to start with the most complicated molecules (those containing the greatest number of atoms). For example, consider the reaction of ethanol with oxygen, given by the unbalanced equation

\[ C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g) \]

which can be represented by the following molecular models:

Notice that the carbon and hydrogen atoms are not balanced. There are two carbon atoms on the left and one on the right, and there are six hydrogens on the left and two on the right. We need to find the correct numbers of reactants and products so that we have the same number of all types of atoms among the reactants and products. We will balance the equation “by inspection” (a systematic trial-and-error procedure).

The most complicated molecule here is \( C_2H_5OH \). We will begin by balancing the products that contain the atoms in \( C_2H_5OH \). Since \( C_2H_5OH \) contains two carbon atoms, we place the coefficient 2 before the \( CO_2 \) to balance the carbon atoms:

\[ C_2H_5OH(l) + O_2(g) \rightarrow 2CO_2(g) + H_2O(g) \]

Since \( C_2H_5OH \) contains six hydrogen atoms, the hydrogen atoms can be balanced by placing a 3 before the \( H_2O \):
Last, we balance the oxygen atoms. Note that the right side of the preceding equation contains seven oxygen atoms, whereas the left side has only three. We can correct this by putting a 3 before the $O_2$ to produce the balanced equation:

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

Now we check:

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

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>2 C atoms</td>
<td>2 C atoms</td>
</tr>
<tr>
<td>6 H atoms</td>
<td>6 H atoms</td>
</tr>
<tr>
<td>7 O atoms</td>
<td>7 O atoms</td>
</tr>
</tbody>
</table>

The equation is balanced.

The balanced equation can be represented as follows:

You can see that all the elements balance.

**Problem-Solving Strategy**

**Writing and Balancing the Equation for a Chemical Reaction**

1. Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?

2. Write the *unbalanced* equation that summarizes the reaction described in Step 1.

3. Balance the equation by inspection, starting with the most complicated molecule(s). Determine what coefficients are necessary so that the same number of each type of atom appears on both reactant and product sides. Do not change the identities (formulas) of any of the reactants or products.
Critical Thinking

One part of the problem-solving strategy for balancing chemical equations is “starting with the most complicated molecule.” What if you started with a different molecule? Could you still eventually balance the chemical equation? How would this approach be different from the suggested technique?

Interactive Example 3.13

Balancing a Chemical Equation I

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, \((\text{NH}_4)_2\text{Cr}_2\text{O}_7\), a vivid orange compound, is ignited, a spectacular reaction occurs, as shown in the two photographs. Although the reaction is actually somewhat more complex, let’s assume here that the products are solid chromium(III) oxide, nitrogen gas (consisting of \(\text{N}_2\) molecules), and water vapor. Balance the equation for this reaction.

Solution

1. From the description given, the reactant is solid ammonium dichromate, \((\text{NH}_4)_2\text{Cr}_2\text{O}_7\), and the products are nitrogen gas, \(\text{N}_2\), water vapor, \(\text{H}_2\text{O}\), and solid chromium(III) oxide, \(\text{Cr}_2\text{O}_3\). The formula for chromium(III) oxide can be determined by recognizing that the Roman numeral III means that \(\text{Cr}^{3+}\) ions are present. For a neutral compound, the formula must then be \(\text{Cr}_2\text{O}_3\), since each oxide ion is \(\text{O}^{2-}\).

2. The unbalanced equation is

\[
(\text{NH}_4)_2\text{Cr}_2\text{O}_7(s) \rightarrow \text{Cr}_2\text{O}_3(s) + \text{N}_2(g) + \text{H}_2\text{O}(g)
\]

3. Note that nitrogen and chromium are balanced (two nitrogen atoms and two chromium atoms on each side), but hydrogen and oxygen are not. A coefficient of 4 for \(\text{H}_2\text{O}\) balances the hydrogen atoms:

\[
(\text{NH}_4)_2\text{Cr}_2\text{O}_7(s) \rightarrow \text{Cr}_2\text{O}_3(s) + \text{N}_2(g) + 4\text{H}_2\text{O}(g)
\]

\[
(4 \times 2)\text{H} \quad \text{and} \quad (4 \times 2)\text{H}
\]

Note that in balancing the hydrogen we also have balanced the oxygen, since there are seven oxygen atoms in the reactants and in the products.
Reality Check

\[
\begin{align*}
2 \text{ N}, & \quad 8 \text{ H}, \quad 2 \text{ Cr}, \quad 7 \text{ O} & \rightarrow & \quad 2 \text{ N}, \quad 8 \text{ H}, \quad 2 \text{ Cr}, \quad 7 \text{ O} \\
\text{Reactant atoms} & & & \text{Product atoms}
\end{align*}
\]

The equation is balanced.

Decomposition of ammonium dichromate.

Photos: Ken O’Donoghue © Cengage Learning

See Exercises 3.95 through 3.98

Interactive Example 3.14

Balancing a Chemical Equation II

At 1000°C, ammonia gas, \( \text{NH}_3(g) \), reacts with oxygen gas to form gaseous nitric oxide, \( \text{NO}(g) \), and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

Solution

1. The unbalanced equation for the reaction is

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

2. Because all the molecules in this equation are of about equal complexity, where we start in balancing it is rather arbitrary. Let's begin by balancing the hydrogen. A coefficient of \( 2 \) for \( \text{NH}_3 \) and a coefficient of \( 3 \) for \( \text{H}_2\text{O} \) give six atoms of hydrogen on both sides:
The nitrogen can be balanced with a coefficient of 2 for NO:

$$2\text{NH}_3(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) + 3\text{H}_2\text{O}(g)$$

Finally, note that there are two atoms of oxygen on the left and five on the right. The oxygen can be balanced with a coefficient of \( \frac{5}{2} \) for \( \text{O}_2 \):

$$2\text{NH}_3(g) + \frac{5}{2} \text{O}_2(g) \rightarrow 2\text{NO}(g) + 3\text{H}_2\text{O}(g)$$

However, the usual custom is to have whole-number coefficients. We simply multiply the entire equation by 2:

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

**Reality Check**

There are 4\text{N}, 12\text{H}, and 10\text{O} on both sides, so the equation is balanced.

We can represent this balanced equation visually as

See Exercises 3.99 through 3.104

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**3.10 Stoichiometric Calculations: Amounts of Reactants and Products**

**AP Learning Objectives**

- LO 1.17 The student is able to express the law of conservation of mass

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Chapter 3: Stoichiometry: 3.10 Stoichiometric Calculations: Amounts of Reactants and Products
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quantitatively and qualitatively using symbolic representations and particulate drawings. [See SP 1.5; Essential knowledge 1.E.1]

LO 1.19 The student can design, and/or interpret data from, an experiment that uses gravimetric analysis to determine the concentration of an analyte in a solution. [See SP 4.2, 5.1; Essential knowledge 1.E.2] (see APEC Lab 3 “Gravimetric Analysis of a Sulfate Mixture”)

As we have seen in previous sections of this chapter, the coefficients in chemical equations represent numbers of molecules, not masses of molecules. However, when a reaction is to be run in a laboratory or chemical plant, the amounts of substances needed cannot be determined by counting molecules directly. Counting is always done by weighing. In this section we will see how chemical equations can be used to determine the masses of reacting chemicals.

Chemical connections

**High Mountains—Low Octane**

The next time you visit a gas station, take a moment to note the octane rating that accompanies the grade of gasoline that you are purchasing. The gasoline is priced according to its octane rating—a measure of the fuel’s antiknock properties. In a conventional internal combustion engine, gasoline vapors and air are drawn into the combustion cylinder on the downward stroke of the piston. This air–fuel mixture is compressed on the upward piston stroke (compression stroke), and a spark from the sparkplug ignites the mix. The rhythmic combustion of the air–fuel mix occurring sequentially in several cylinders furnishes the power to propel the vehicle down the road. Excessive heat and pressure (or poor-quality fuel) within the cylinder may cause the premature combustion of the mixture—commonly known as engine “knock” or “ping.” Over time, this engine knock can damage the engine, resulting in inefficient performance and costly repairs.

A consumer typically is faced with three choices of gasoline, with octane ratings of 87 (regular), 89 (midgrade), and 93 (premium). But if you happen to travel or live in the higher elevations of the Rocky Mountain states, you might be surprised to find different octane ratings at the gasoline pumps. The reason for this provides a lesson in stoichiometry. At higher elevations the air is less dense—the volume of oxygen per unit volume of air is smaller. Most engines are designed to achieve a 14:1 oxygen-to-fuel ratio in the cylinder prior to combustion. If less oxygen is available, then less fuel is required to achieve this optimal ratio. In turn, the lower volumes of oxygen and fuel result in a lower pressure in the cylinder. Because high pressure tends to promote knocking, the lower pressure within engine cylinders at higher elevations promotes a more controlled combustion of the air–fuel mixture, and therefore octane requirements are lower. While consumers in the Rocky Mountain
states can purchase three grades of gasoline, the octane ratings of these fuel blends are different from those in the rest of the United States. In Denver, Colorado, regular gasoline is 85 octane, midgrade is 87 octane, and premium is 91 octane—2 points lower than gasoline sold in most of the rest of the country.

To develop the principles for dealing with the stoichiometry of reactions, we will consider the reaction of propane with oxygen to produce carbon dioxide and water. We will consider the question: “What mass of oxygen will react with 96.1 g of propane?” In doing stoichiometry, the first thing we must do is write the balanced chemical equation for the reaction. In this case the balanced equation is

\[
C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)
\]

which can be visualized as

![Chemical reaction diagram](image)

This equation means that 1 mole of \( C_3H_8 \) reacts with 5 moles of \( O_2 \) to produce 3 moles of \( CO_2 \) and 4 moles of \( H_2O \). To use this equation to find the masses of reactants and products, we must be able to convert between masses and moles of substances. Thus we must first ask: “How many moles of propane are present in 96.1 g of propane?” The molar mass of propane to three significant figures is 44.1 (that is, \( 3 \times 12.01 + 8 \times 1.008 \)). The moles of propane can be calculated as follows:

\[
96.1 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.1 \text{ g } C_3H_8} = 2.18 \text{ mol } C_3H_8
\]

Next we must take into account the fact that each mole of propane reacts with 5 moles of oxygen. The best way to do this is to use the balanced equation to construct a mole ratio (the ratio of moles of one substance to moles of another substance in a balanced chemical equation). In this case we want to convert from moles of propane to moles of oxygen. From the balanced equation, we see that 5 moles of \( O_2 \) are required for each mole of \( C_3H_8 \), so the appropriate ratio is

\[
\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}
\]

Multiplying the number of moles of \( C_3H_8 \) by this factor gives the number of moles of \( O_2 \) required:
Notice that the mole ratio is set up so that the moles of \( \text{C}_3\text{H}_8 \) cancel out, and the units that result are moles of \( \text{O}_2 \).

Since the original question asked for the mass of oxygen needed to react with 96.1g of propane, the 10.9 moles of \( \text{O}_2 \) must be converted to grams. Since the molar mass of \( \text{O}_2 \) is 32.0g/mol,

\[
10.9 \text{ mol O}_2 \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = 349 \text{ g O}_2
\]

Therefore, 349g of oxygen are required to burn 96.1g of propane.

This example can be extended by asking: “What mass of carbon dioxide is produced when 96.1g of propane are combusted with oxygen?” In this case we must convert between moles of propane and moles of carbon dioxide. This can be accomplished by looking at the balanced equation, which shows that 3 moles of \( \text{CO}_2 \) are produced for each mole of \( \text{C}_3\text{H}_8 \) reacted. The mole ratio needed to convert from moles of propane to moles of carbon dioxide is

\[
\frac{3\text{ mol CO}_2}{1\text{ mol C}_3\text{H}_8}
\]

The conversion is

\[
2.18 \text{ mol C}_3\text{H}_8 \times \frac{3\text{ mol CO}_2}{1\text{ mol C}_3\text{H}_8} = 6.54\text{ mol CO}_2
\]

Then, using the molar mass of \( \text{CO}_2 \) (44.0g/mol), we calculate the mass of \( \text{CO}_2 \) produced:

\[
6.54 \text{ mol CO}_2 \times \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} = 288 \text{ g CO}_2
\]

We will now summarize the sequence of steps needed to carry out stoichiometric calculations.

96.1 g \( \text{C}_3\text{H}_8 \) \( \frac{1 \text{ mol C}_3\text{H}_8}{44.1 \text{ g C}_3\text{H}_8} \) 2.18 mol \( \text{C}_3\text{H}_8 \) \( \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \) 6.54 mol \( \text{CO}_2 \)

\[
\frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} \rightarrow 288 \text{ g CO}_2
\]

Critical Thinking
Your lab partner has made the observation that you always take the mass of chemicals in lab, but then you use mole ratios to balance the equation. “Why not use the masses in the equation?” your partner asks. What if your lab partner decided to balance equations by using masses as coefficients? Is this even possible? Why or why not?

Problem-Solving Strategy

**Calculating Masses of Reactants and Products in Chemical Reactions**

1. Balance the equation for the reaction.

2. Convert the known mass of the reactant or product to moles of that substance.

3. Use the balanced equation to set up the appropriate mole ratios.

4. Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.

5. Convert from moles back to grams if required by the problem.

Interactive Example 3.15

**Chemical Stoichiometry I**

Solid lithium hydroxide is used in space vehicles to remove exhaled
carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

**Solution**

**Where are we going?**

To find the mass of \( \text{CO}_2 \) absorbed by 1.00 kg LiOH

**What do we know?**

- Chemical reaction

\[
\text{LiOH}(s) + \text{CO}_2(g) \rightarrow \text{Li}_2\text{CO}_3(s) + \text{H}_2\text{O}(l)
\]

- 1.00 kg LiOH

**What information do we need to find the mass of \( \text{CO}_2 \)?**

- Balanced equation for the reaction

**How do we get there?**

1. **What is the balanced equation?**

\[
2\text{LiOH}(s) + \text{CO}_2(g) \rightarrow \text{Li}_2\text{CO}_3(s) + \text{H}_2\text{O}(l)
\]

2. **What are the moles of LiOH?**

To find the moles of LiOH, we need to know the molar mass.

**What is the molar mass for LiOH?**

\[
6.941 + 16.00 + 1.008 = 23.95 \text{ g} / \text{mol}
\]

Now we use the molar mass to find the moles of LiOH:

\[
1.00 \text{ kg LiOH} \times \frac{1000 \text{ g LiOH}}{1 \text{ kg LiOH}} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} = 41.8 \text{ mol LiOH}
\]

3. **What is the mole ratio between \( \text{CO}_2 \) and LiOH in the balanced equation?**

\[
\frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}}
\]

4. **What are the moles of \( \text{CO}_2 \)?**
5. What is the mass of CO₂ formed from 1.00 kg LiOH?

\[
41.8 \text{ mol LiOH} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} = 20.9 \text{ mol CO}_2
\]

\[
20.9 \text{ mol CO}_2 \times \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} = 9.20 \times 10^2 \text{ g CO}_2
\]

- Thus 920 g of CO₂(g) will be absorbed by 1.00 kg of LiOH(s).

See Exercises 3.105 and 3.106

Interactive Example 3.16

Chemical Stoichiometry II

Baking soda (NaHCO₃) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach:

\[ \text{NaHCO}_3(s) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(aq) \]

Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, is also used as an antacid:

\[ \text{Mg(OH)}_2(s) + 2\text{HCl}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{MgCl}_2(aq) \]

Which is the more effective antacid per gram, NaHCO₃ or Mg(OH)₂?

Solution

Where are we going?

To compare the acid neutralizing power of NaHCO₃ and Mg(OH)₂ per gram

What do we know?

- Balanced equations for the reactions
- 1.00 g NaHCO₃
- 1.00 g Mg(OH)₂

How do we get there?

For NaHCO₃

\[
41.8 \text{ mol LiOH} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} = 20.9 \text{ mol CO}_2
\]

\[
20.9 \text{ mol CO}_2 \times \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} = 9.20 \times 10^2 \text{ g CO}_2
\]
1. What is the balanced equation?

\[ \text{NaHCO}_3(s) + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2(aq) \]

2. What are the moles of NaHCO\(_3\) in 1.00g?

To find the moles of NaHCO\(_3\), we need to know the molar mass (84.01 g/mol).

\[
1.00 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} = 1.19 \times 10^{-2} \text{ mol NaHCO}_3
\]

3. What is the mole ratio between HCl and NaHCO\(_3\) in the balanced equation?

\[
\frac{1 \text{ mol HCl}}{1 \text{ mol NaHCO}_3}
\]

4. What are the moles of HCl?

\[
1.19 \times 10^{-2} \text{ mol NaHCO}_3 \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaHCO}_3} = 1.19 \times 10^{-2} \text{ mol HCl}
\]

Thus 1.00g of NaHCO\(_3\) will neutralize \(1.19 \times 10^{-2}\) mole of HCl.

Milk of magnesia contains a suspension of \(\text{Mg(OH)}_2(s)\).
2. What are the moles of $\text{Mg(OH)}_2$ in 1.00 g?

To find the moles of $\text{Mg(OH)}_2$, we need to know the molar mass (58.32 g/mol).

$$1.00 \text{ g } \text{Mg(OH)}_2 \times \frac{1 \text{ mol } \text{Mg(OH)}_2}{58.32 \text{ g } \text{Mg(OH)}_2} = 1.71 \times 10^{-2} \text{ mol } \text{Mg(OH)}_2$$

3. What is the mole ratio between $\text{HCl}$ and $\text{Mg(OH)}_2$ in the balanced equation?

$$\frac{2 \text{ mol HCl}}{1 \text{ mol } \text{Mg(OH)}_2}$$

4. What are the moles of $\text{HCl}$?

$$1.71 \times 10^{-2} \text{ mol } \text{Mg(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } \text{Mg(OH)}_2} = 3.42 \times 10^{-2} \text{ mol HCl}$$

Thus 1.00 g of $\text{Mg(OH)}_2$ will neutralize $3.42 \times 10^{-2}$ mole of $\text{HCl}$.

- Since 1.00 g $\text{NaHCO}_3$ neutralizes $1.19 \times 10^{-2}$ mole of $\text{HCl}$ and 1.00 g $\text{Mg(OH)}_2$ neutralizes $3.42 \times 10^{-2}$ mole of $\text{HCl}$, $\text{Mg(OH)}_2$ is the more effective antacid.

See Exercises 3.107 and 3.108

---

**3.11 The Concept of Limiting Reactant**

AP Learning Objectives

- **LO 1.17** The student is able to express the law of conservation of mass quantitatively and qualitatively using symbolic representations and particulate drawings. [See SP 1.5; Essential knowledge 1.E.1]

- **LO 3.3** The student is able to use stoichiometric calculations to predict the results of performing a reaction in the laboratory and/or to analyze deviations from the expected results. [See SP 2.2, 5.1; Essential knowledge 3.A.2]

- **LO 3.4** The student is able to relate quantities (measured mass of $\text{Mg(OH)}_2$ in 1.00 g).
Suppose you have a part-time job in a sandwich shop. One very popular sandwich is always made as follows:

\[ 2 \text{ slices bread} + 3 \text{slices meat} + 1 \text{ slice cheese} \rightarrow \text{ sandwich} \]

Assume that you come to work one day and find the following quantities of ingredients:

- 8 slices bread
- 9 slices meat
- 5 slices cheese

How many sandwiches can you make? What will be left over?

To solve this problem, let’s see how many sandwiches we can make with each component:

<table>
<thead>
<tr>
<th>Component</th>
<th>Quantity</th>
<th>Sandwiches</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bread</td>
<td>8 slices</td>
<td>4</td>
</tr>
<tr>
<td>Meat</td>
<td>9 slices</td>
<td>3</td>
</tr>
<tr>
<td>Cheese</td>
<td>5 slices</td>
<td>5</td>
</tr>
</tbody>
</table>

How many sandwiches can you make? The answer is three. When you run out of meat, you must stop making sandwiches. The meat is the limiting ingredient (Fig. 3.9).

**Figure 3.9**

Making sandwiches.
What do you have left over? Making three sandwiches requires six pieces of bread. You started with eight slices, so you have two slices of bread left. You also used three pieces of cheese for the three sandwiches, so you have two pieces of cheese left.

In this example, the ingredient present in the largest number (the meat) was actually the component that limited the number of sandwiches you could make. This situation arose because each sandwich required three slices of meat—more than the quantity required of any other ingredient.

When molecules react with each other to form products, considerations very similar to those involved in making sandwiches arise. We can illustrate these ideas with the reaction of $\text{N}_2(g)$ and $\text{H}_2(g)$ to form $\text{NH}_3(g)$:

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

Consider the following container of $\text{N}_2(g)$ and $\text{H}_2(g)$:
What will this container look like if the reaction between \( \text{N}_2 \) and \( \text{H}_2 \) proceeds to completion? To answer this question, you need to remember that each \( \text{N}_2 \) requires \( 3 \text{H}_2 \) molecules to form \( 2 \text{NH}_3 \). To make things clear, we will circle groups of reactants:

In this case, the mixture of \( \text{N}_2 \) and \( \text{H}_2 \) contained just the number of molecules needed to form \( \text{NH}_3 \) with nothing left over. That is, the ratio of the number of \( \text{H}_2 \) molecules to \( \text{N}_2 \) molecules was

\[
\frac{15 \text{H}_2}{5 \text{N}_2} = \frac{3 \text{H}_2}{1 \text{N}_2}
\]

This ratio exactly matches the numbers in the balanced equation

\[
3 \text{H}_2(g) + \text{N}_2(g) \rightarrow 2 \text{NH}_3(g)
\]

This type of mixture is called a **stoichiometric mixture**—one that contains the relative amounts of reactants that match the numbers in the balanced equation. In this case all reactants will be consumed to form products.

Now consider another container of \( \text{N}_2(g) \) and \( \text{H}_2(g) \):
What will the container look like if the reaction between $\text{N}_2(g)$ and $\text{H}_2(g)$ proceeds to completion? Remember that each $\text{N}_2$ requires $3\text{H}_2$. Circling groups of reactants, we have

In this case, the hydrogen ($\text{H}_2$) is limiting. That is, the $\text{H}_2$ molecules are used up before all the $\text{N}_2$ molecules are consumed. In this situation the amount of hydrogen limits the amount of product (ammonia) that can form—hydrogen is the limiting reactant. Some $\text{N}_2$ molecules are left over in this case because the reaction runs out of $\text{H}_2$ molecules first. To determine how much product can be formed from a given mixture of reactants, we have to look for the reactant that is limiting—the one that runs out first and thus limits the amount of product that can form. In some cases, the mixture of reactants might be stoichiometric—that is, all reactants run out at the same time. In general, however, you cannot assume that a given mixture of reactants is a stoichiometric mixture, so you must determine whether one of the reactants is limiting. The reactant that runs out first and thus limits the amounts of products that can form is called the **limiting reactant** (the reactant that is completely consumed when a reaction is run to completion).

To this point we have considered examples where the numbers of reactant molecules could be counted. In “real life” you can’t count the molecules directly—you can’t see them, and even if you could, there would be far too many to count. Instead, you must count by weighing. We must therefore explore how to find the limiting reactant, given the masses of the reactants.
A. Determination of Limiting Reactant Using Reactant Quantities

There are two ways to determine the limiting reactant in a chemical reaction. One involves comparing the moles of reactants to see which runs out first. We will consider this approach here.

In the laboratory or chemical plant, we work with much larger quantities than the few molecules of the preceding example. Therefore, we must learn to deal with limiting reactants using moles. The ideas are exactly the same, except that we are using moles of molecules instead of individual molecules. For example, suppose 25.0 kg of nitrogen and 5.00 kg of hydrogen are mixed and reacted to form ammonia. How do we calculate the mass of ammonia produced when this reaction is run to completion (until one of the reactants is completely consumed)?

As in the preceding example, we must use the balanced equation

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

to determine whether nitrogen or hydrogen is the limiting reactant and then to determine the amount of ammonia that is formed. We first calculate the moles of reactants present:

\[
\begin{align*}
25.0 \text{ kg N}_2 & \times \frac{1000 \text{ g N}_2}{1 \text{ kg N}_2} \times \frac{1 \text{ mol N}_2}{28.0 \text{ g N}_2} = 8.93 \times 10^2 \text{ mol N}_2 \\
5.00 \text{ kg H}_2 & \times \frac{1000 \text{ g H}_2}{1 \text{ kg H}_2} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 2.48 \times 10^3 \text{ mol H}_2
\end{align*}
\]

Since 1 mole of N$_2$ reacts with 3 moles of H$_2$, the number of moles of H$_2$ that will react exactly with $8.93 \times 10^2$ moles of N$_2$ is

\[
8.93 \times 10^2 \frac{\text{mol N}_2}{1 \text{ mol N}_2} \times 3 \text{ mol H}_2 = 2.68 \times 10^3 \text{ mol H}_2
\]

Thus $8.93 \times 10^2$ moles of N$_2$ requires $2.68 \times 10^3$ moles of H$_2$ to react completely. However, in this case, only $2.48 \times 10^3$ moles of H$_2$ are present. This means that the hydrogen will be consumed before the nitrogen. Thus hydrogen is the limiting reactant in this particular situation, and we must use the amount of hydrogen to compute the quantity of ammonia formed:

\[
2.48 \times 10^3 \frac{\text{mol H}_2}{1 \text{ mol H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 1.65 \times 10^3 \text{ mol NH}_3
\]

Converting moles to kilograms gives

\[
1.65 \times 10^3 \frac{\text{mol NH}_3}{1 \text{ mol NH}_3} \times \frac{17.0 \text{ g NH}_3}{1 \text{ mol NH}_3} = 2.80 \times 10^4 \text{ g NH}_3 = 28.0 \text{ kg NH}_3
\]
Note that to determine the limiting reactant, we could have started instead with the given amount of hydrogen and calculated the moles of nitrogen required:

\[
2.48 \times 10^3 \text{ mol H}_2 \times \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} = 8.27 \times 10^2 \text{ mol N}_2
\]

Thus \(2.48 \times 10^3\) moles of \(\text{H}_2\) requires \(8.27 \times 10^2\) moles of \(\text{N}_2\). Since \(8.93 \times 10^2\) moles of \(\text{N}_2\) are actually present, the nitrogen is in excess. The hydrogen will run out first, and thus again we find that hydrogen limits the amount of ammonia formed.

A related but simpler way to determine which reactant is limiting is to compare the mole ratio of the substances required by the balanced equation with the mole ratio of reactants actually present. For example, in this case the mole ratio of \(\text{H}_2\) to \(\text{N}_2\) required by the balanced equation is

\[
\frac{3\text{ mol H}_2}{1\text{ mol N}_2}
\]

That is,

\[
\frac{\text{mol H}_2}{\text{mol N}_2} \text{ (required)} = \frac{3}{1} = 3
\]

In this experiment we have \(2.48 \times 10^3\) moles of \(\text{H}_2\) and \(8.93 \times 10^2\) moles of \(\text{N}_2\). Thus the ratio

\[
\frac{\text{mol H}_2}{\text{mol N}_2} \text{ (actual)} = \frac{2.48 \times 10^3}{8.93 \times 10^2} = 2.78
\]

Since 2.78 is less than 3, the actual mole ratio of \(\text{H}_2\) to \(\text{N}_2\) is too small, and \(\text{H}_2\) must be limiting. If the actual \(\text{H}_2\) to \(\text{N}_2\) mole ratio had been greater than 3, then the \(\text{H}_2\) would have been in excess and the \(\text{N}_2\) would be limiting.

B. Determination of Limiting Reactant Using Quantities of Products Formed

A second method for determining which reactant in a chemical reaction is limiting is to consider the amounts of products that can be formed by completely consuming each reactant. The reactant that produces the smallest amount of product must run out first and thus be limiting. To see how this works, consider again the reaction of \(25.0\text{kg}(8.93 \times 10^2\text{ moles})\) of nitrogen with \(5.00\text{kg}(2.48 \times 10^3\text{ moles})\) of hydrogen.

We will now use these amounts of reactants to determine how much \(\text{NH}_3\) would form. Since \(1\text{ mole of N}_2\) forms \(2\text{ moles of NH}_3\), the amount of \(\text{NH}_3\) that would be formed if all of the \(\text{N}_2\) was used up is calculated as follows:

\[
8.93 \times 10^2 \text{ mol N}_2 \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} = 1.79 \times 10^3 \text{ mol NH}_3
\]

Next we will calculate how much \(\text{NH}_3\) would be formed if the \(\text{H}_2\) was completely used up:
2.48 \times 10^3 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 1.65 \times 10^3 \text{ mol NH}_3

Because a smaller amount of \text{NH}_3 is produced from the \text{H}_2 than from the \text{N}_2, the amount of \text{H}_2 must be limiting.

Thus because the \text{H}_2 is the limiting reactant, the amount of \text{NH}_3 that can form is 1.65 \times 10^3 \text{ moles}. Converting moles to kilograms gives:

\[1.65 \times 10^3 \text{ mol NH}_3 \times \frac{17 \text{ g NH}_3}{1 \text{ mol NH}_3} = 2.80 \times 10^4 \text{ g NH}_3 = 28.0 \text{ kg NH}_3\]

Interactive Example 3.17

**Stoichiometry: Limiting Reactant**

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1g of \text{NH}_3 is reacted with 90.4g of \text{CuO}, which is the limiting reactant? How many grams of \text{N}_2 will be formed?

**Solution**

**Where are we going?**

To find the limiting reactant

To find the mass of \text{N}_2 produced

**What do we know?**

- The chemical reaction

\[\text{NH}_3(g) + \text{CuO}(s) \rightarrow \text{N}_2(g) + \text{Cu}(s) + \text{H}_2\text{O}(g)\]

- 18.1g \text{NH}_3

- 90.4g \text{CuO}

**What information do we need?**

- Balanced equation for the reaction

- Moles of \text{NH}_3

- Moles of \text{CuO}
How do we get there?

To find the limiting reactant

What is the balanced equation?

\[ 2\text{NH}_3(g) + 3\text{CuO}(s) \rightarrow \text{N}_2(g) + 3\text{Cu}(s) + 3\text{H}_2\text{O}(g) \]

What are the moles of \text{NH}_3 and \text{CuO}?

To find the moles, we need to know the molar masses.

\[
\begin{align*}
\text{NH}_3 & \quad 17.03 \text{ g/mol} \\
\text{CuO} & \quad 79.55 \text{ g/mol}
\end{align*}
\]

\[
\begin{align*}
18.1 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} & = 1.06 \text{ mol NH}_3 \\
90.4 \text{ g CuO} \times \frac{1 \text{ mol CuO}}{79.55 \text{ g CuO}} & = 1.14 \text{ mol CuO}
\end{align*}
\]

A. First we will determine the limiting reactant by comparing the moles of reactants to see which one is consumed first.

What is the mole ratio between \text{NH}_3 and \text{CuO} in the balanced equation?

\[
\frac{3 \text{ mol CuO}}{2 \text{ mol NH}_3}
\]

How many moles of \text{CuO} are required to react with 1.06 moles of \text{NH}_3?

\[
1.06 \text{ mol NH}_3 \times \frac{3 \text{ mol CuO}}{2 \text{ mol NH}_3} = 1.59 \text{ mol CuO}
\]

- Thus 1.59 moles of \text{CuO} are required to react with 1.06 moles of \text{NH}_3. Since only 1.14 moles of \text{CuO} are actually present, the amount of \text{CuO} is limiting; \text{CuO} will run out before \text{NH}_3 does. We can verify this conclusion by comparing the mole ratio of \text{CuO} and \text{NH}_3 required by the balanced equation:

\[
\frac{\text{mol CuO}}{\text{mol NH}_3} \text{ (required)} = \frac{3}{2} = 1.5
\]

with the mole ratio actually present:

\[
\frac{\text{mol CuO}}{\text{mol NH}_3} \text{ (actual)} = \frac{1.14}{1.06} = 1.08
\]

- Since the actual ratio is too small (less than 1.5), \text{CuO} is the limiting reactant.
B. Alternatively we can determine the limiting reactant by computing the moles of $N_2$ that would be formed by complete consumption of $\text{NH}_3$ and $\text{CuO}$:

\[
1.06 \text{ mol } \text{NH}_3 \times \frac{1\text{ mol } N_2}{2\text{ mol } \text{NH}_3} = 0.530\text{ mol } N_2 \\
1.14 \text{ mol } \text{CuO} \times \frac{1\text{ mol } N_2}{3\text{ mol } \text{CuO}} = 0.380\text{ mol } N_2
\]

As before, we see that the $\text{CuO}$ is limiting since it produces the smaller amount of $N_2$.

To find the mass of $N_2$ produced

*What are the moles of $N_2$ formed?*

Because $\text{CuO}$ is the limiting reactant, we must use the amount of $\text{CuO}$ to calculate the amount of $N_2$ formed.

*What is the mole ratio between $N_2$ and $\text{CuO}$ in the balanced equation?*

\[
\frac{1\text{ mol } N_2}{3\text{ mol } \text{CuO}}
\]

*What are the moles of $N_2$?*

\[
1.14 \text{ mol } \text{CuO} \times \frac{1\text{ mol } N_2}{3\text{ mol } \text{CuO}} = 0.380\text{ mol } N_2
\]

*What mass of $N_2$ is produced?*

Using the molar mass of $N_2$ (28.02 g/mol), we can calculate the mass of $N_2$ produced:

\[
0.380 \text{ mol } N_2 \times \frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2} = 10.6 \text{ g } N_2
\]

See Exercises 3.117 through 3.122

The amount of a product formed when the limiting reactant is completely consumed is called the **theoretical yield** (the maximum amount of a given product that can be formed when the limiting reactant is completely consumed) of that product. In Example 3.17, 10.6 g of nitrogen represent the theoretical yield. This is the maximum amount of nitrogen that can be produced from the quantities of reactants used. Actually, the amount of product predicted by the theoretical yield is seldom obtained because of side reactions (other reactions that involve one or more of the reactants or products) and other complications. The **actual yield** of product is often given as a percentage of the theoretical yield. This is called the **percent yield** (the actual yield of a product as a percentage of the theoretical yield):
For example, if the reaction considered in Example 3.17 actually gave 6.63 g of nitrogen instead of the predicted 10.6 g, the percent yield of nitrogen would be:

\[
\frac{6.63 \text{ g } \text{N}_2}{10.6 \text{ g } \text{N}_2} \times 100\% = 62.5\%
\]

Interactive Example 3.18

Calculating Percent Yield

Methanol (\(\text{CH}_3\text{OH}\)), also called methyl alcohol, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combining gaseous carbon monoxide and hydrogen. Suppose 68.5 kg \(\text{CO}(g)\) is reacted with 8.60 kg \(\text{H}_2(g)\). Calculate the theoretical yield of methanol. If \(3.57 \times 10^4\text{ g} \text{ CH}_3\text{OH}\) is actually produced, what is the percent yield of methanol?

Solution

Where are we going?

To calculate the theoretical yield of methanol

To calculate the percent yield of methanol

What do we know?

- The chemical reaction
  \[\text{H}_2(g) + \text{CO}(g) \rightarrow \text{CH}_3\text{OH}(l)\]
- 68.5 kg \(\text{CO}(g)\)
- 8.60 kg \(\text{H}_2(g)\)
- \(3.57 \times 10^4\text{ g} \text{ CH}_3\text{OH}\) is produced
What information do we need?

- Balanced equation for the reaction
- Moles of H₂
- Moles of CO
- Which reactant is limiting
- Amount of CH₃OH produced

How do we get there?

To find the limiting reactant

What is the balanced equation?

\[ 2\text{H}_2(g) + \text{CO}(g) \rightarrow \text{CH}_3\text{OH}(l) \]

What are the moles of H₂ and CO?

To find the moles, we need to know the molar masses.

\[
\begin{align*}
\text{H}_2 & \quad 2.016 \text{g/mol} \\
\text{CO} & \quad 28.02 \text{g/mol}
\end{align*}
\]

\[
\begin{align*}
68.5 \text{ kg CO} & \times \frac{1000 \text{ g CO}}{1 \text{ kg CO}} \times \frac{1 \text{ mol CO}}{28.02 \text{ g CO}} = 2.44 \times 10^3 \text{ mol CO} \\
8.60 \text{ kg H}_2 & \times \frac{1000 \text{ g H}_2}{1 \text{ kg H}_2} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 4.27 \times 10^3 \text{ mol H}_2
\end{align*}
\]

A. Determination of Limiting Reactant Using Reactant Quantities

What is the mole ratio between H₂ and CO in the balanced equation?

\[
\frac{2\text{mol H}_2}{1\text{mol CO}}
\]

How does the actual mole ratio compare to the stoichiometric ratio?

To determine which reactant is limiting, we compare the mole ratio of H₂ and CO required by the balanced equation

\[
\frac{\text{mol H}_2}{\text{mol CO}} \text{ (required)} = \frac{2}{1} = 2
\]

with the actual mole ratio
Since the actual mole ratio of $\text{H}_2$ to CO is smaller than the required ratio, $\text{H}_2$ is limiting.

B. Determination of Limiting Reactant Using Quantities of Products Formed

We can also determine the limiting reactant by calculating the amounts of $\text{CH}_3\text{OH}$ formed by complete consumption of $\text{CO}(g)$ and $\text{H}_2(g)$:

\[
2.44 \times 10^3 \text{ mol CO} \times \frac{1 \text{ mol CH}_3\text{OH}}{1 \text{ mol CO}} = 2.44 \times 10^3 \text{ mol CH}_3\text{OH}
\]

\[
4.27 \times 10^3 \text{ mol H}_2 \times \frac{1 \text{ mol CH}_3\text{OH}}{2 \text{ mol H}_2} = 2.14 \times 10^3 \text{ mol CH}_3\text{OH}
\]

Since complete consumption of the $\text{H}_2$ produces the smaller amount of $\text{CH}_3\text{OH}$, the $\text{H}_2$ is the limiting reactant as we determined above.

To calculate the theoretical yield of methanol

*What are the moles of $\text{CH}_3\text{OH}$ formed?*

We must use the amount of $\text{H}_2$ and the mole ratio between $\text{H}_2$ and $\text{CH}_3\text{OH}$ to determine the maximum amount of methanol that can be produced:

\[
4.27 \times 10^3 \text{ mol H}_2 \times \frac{1 \text{ mol CH}_3\text{OH}}{2 \text{ mol H}_2} = 2.14 \times 10^3 \text{ mol CH}_3\text{OH}
\]

*What is the theoretical yield of $\text{CH}_3\text{OH}$ in grams?*

\[
2.14 \times 10^3 \text{ mol CH}_3\text{OH} \times \frac{32.04 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} = 6.86 \times 10^4 \text{ g CH}_3\text{OH}
\]

Thus, from the amount of reactants given, the maximum amount of $\text{CH}_3\text{OH}$ that can be formed is $6.86 \times 10^4 \text{ g}$. This is the theoretical yield.

*What is the percent yield of $\text{CH}_3\text{OH}$?*

\[
\text{Percent yield} = \frac{\text{Actual yield (grams)}}{\text{Theoretical yield (grams)}} \times 100 = \frac{3.57 \times 10^4 \text{ g CH}_3\text{OH}}{6.86 \times 10^4 \text{ g CH}_3\text{OH}} \times 100\% = 52.0\%
\]

See Exercises 3.123 and 3.124

**Problem-Solving Strategy**

**Solving a Stoichiometry Problem Involving Masses of Reactants and Products**
1. Write and balance the equation for the reaction.

2. Convert the known masses of substances to moles.

3. Determine which reactant is limiting.

4. Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.

5. Convert from moles to grams, using the molar mass.

This process is summarized in the diagram below:

---

**For Review**

**Key Terms**

- **chemical stoichiometry** (the calculation of the quantities of material consumed and produced in chemical reactions.)

Section 3.2

- **mass spectrometer** (an instrument used to determine the relative masses of atoms by the deflection of their ions on a magnetic field.)

- **average atomic mass** (the weighted average mass of the atoms in a naturally occurring element; also known as atomic weight)

Section 3.3
mole (mol) (the number equal to the number of carbon atoms in exactly 12 grams of pure $^{12}$C: Avogadro's number. One mole represents $6.022 \times 10^{23}$ units.)

Avogadro's number (the number of atoms in exactly 12 grams of pure $^{12}$C, equal to $6.022 \times 10^{23}$.)

Section 3.4

molar mass (the mass in grams of one mole of molecules or formula units of a substance; also called molecular weight.)

Section 3.5

conceptual problem solving ()

Section 3.6

mass percent (the percent by mass of a component of a mixture or of a given element in a compound.)

Section 3.7

empirical formula (the simplest whole-number ratio of atoms in a compound.)

molecular formula (the exact formula of a molecule, giving the types of atoms and the number of each type.)

Section 3.8

chemical equation (a representation of a chemical reaction showing the relative numbers of reactant and product molecules.)

reactants (a starting substance in a chemical reaction. It appears to the left of the arrow in a chemical equation.)

products (a substance resulting from a chemical reaction. It is shown to the right of the arrow in a chemical equation.)

balancing a chemical equation (making sure that all atoms present in the reactants are accounted for among the products)

Section 3.10

mole ratio (the ratio of moles of one substance to moles of another substance in a balanced chemical equation.)

Section 3.11

stoichiometric mixture ()

limiting reactant (the reactant that is completely consumed when a reaction is run to completion.)
theoretical yield (the maximum amount of a given product that can be formed when the limiting reactant is completely consumed.)

percent yield (the actual yield of a product as a percentage of the theoretical yield.)

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For Review

Summary

Stoichiometry

- Deals with the amounts of substances consumed and/or produced in a chemical reaction.
- We count atoms by measuring the mass of the sample.
- To relate mass and the number of atoms, the average atomic mass is required.

Mole

- The amount of carbon atoms in exactly 12g of pure 12C
- 6.022 \times 10^{23} \text{ units of a substance}
- The mass of 1 mole of an element = the atomic mass in grams

Molar Mass

- Mass (g) of 1 mole of a compound or element
- Obtained for a compound by finding the sum of the average masses of its constituent atoms

Percent Composition

- The mass percent of each element in a compound
- \[
  \text{Mass percent} = \frac{\text{mass of element in 1 mole of substance}}{\text{mass of 1 mole of substance}} \times 100%
\]

Empirical Formula

- The simplest whole-number ratio of the various types of atoms in a compound
- Can be obtained from the mass percent of elements in a compound

Molecular Formula
For molecular substances:

- The formula of the constituent molecules
- Always an integer multiple of the empirical formula

For ionic substances:

- The same as the empirical formula

**Chemical Reactions**

- Reactants are turned into products.
- Atoms are neither created nor destroyed.
- All of the atoms present in the reactants must also be present in the products.

**Characteristics of A Chemical Equation**

- Represents a chemical reaction
- Reactants on the left side of the arrow, products on the right side
- When balanced, gives the relative numbers of reactant and product molecules or ions

**Stoichiometry Calculations**

- Amounts of reactants consumed and products formed can be determined from the balanced chemical equation.
- The limiting reactant is the one consumed first, thus limiting the amount of product that can form.

**Yield**

- The theoretical yield is the maximum amount that can be produced from a given amount of the limiting reactant.
- The actual yield, the amount of product actually obtained, is always less than the theoretical yield.

- \[ \text{Percent yield} = \frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\% \]
Review Questions

Answers to the Review Questions can be found on the student website.

1. Explain the concept of “counting by weighing” using marbles as your example.

2. Atomic masses are relative masses. What does this mean?

3. The atomic mass of boron (B) is given in the periodic table as 10.81, yet no single atom of boron has a mass of 10.81u. Explain.

4. What three conversion factors and in what order would you use them to convert the mass of a compound into atoms of a particular element in that compound—for example, from 1.00g aspirin \((C_9H_8O_4)\) to number of hydrogen atoms in the 1.00-g sample?

5. **Fig. 3.5** illustrates a schematic diagram of a combustion device used to analyze organic compounds. Given that a certain amount of a compound containing carbon, hydrogen, and oxygen is combusted in this device, explain how the data relating to the mass of \(CO_2\) produced and the mass of \(H_2O\) produced can be manipulated to determine the empirical formula.

6. What is the difference between the empirical and molecular formulas of a compound? Can they ever be the same? Explain.

7. Consider the hypothetical reaction between \(A_2\) and \(AB\) pictured below.

   ![Reaction Diagram]

   What is the balanced equation? If 2.50 moles of \(A_2\) are reacted with excess \(AB\), what amount (moles) of product will form? If the mass of \(AB\) is 30.0u and the mass of \(A_2\) are 40.0u, what is the mass of the product? If 15.0g of \(AB\) is reacted, what mass of \(A_2\) is required to react with all of the \(AB\), and what mass of product is formed?

8. What is a limiting reactant problem? Explain the method you are going to use to solve limiting reactant problems.

9. Consider the following mixture of \(SO_2(g)\) and \(O_2(g)\).
If SO₂(g) and O₂(g) react to form SO₃(g), draw a representation of the product mixture assuming the reaction goes to completion. What is the limiting reactant in the reaction? If 96.0g of SO₂ react with 32.0gO₂, what mass of product will form?

10. Why is the actual yield of a reaction often less than the theoretical yield?
2. What information do we get from a chemical formula? From a chemical equation?

3. You are making cookies and are missing a key ingredient—eggs. You have most of the other ingredients needed to make the cookies, except you have only 1.33 cups of butter and no eggs. You note that the recipe calls for two cups of butter and three eggs (plus the other ingredients) to make six dozen cookies. You call a friend and have him bring you some eggs.

   a. What number of eggs do you need?
   
   b. If you use all the butter (and get enough eggs), what number of cookies will you make?

Unfortunately, your friend hangs up before you tell him how many eggs you need. When he arrives, he has a surprise for you—to save time, he has broken them all in a bowl for you. You ask him how many he brought, and he replies, “I can’t remember.” You weigh the eggs and find that they weigh 62.1g. Assuming that an average egg weighs 34.21g,

   a. What quantity of butter is needed to react with all the eggs?
   
   b. What number of cookies can you make?
   
   c. Which will you have left over, eggs or butter?
   
   d. What quantity is left over?

4. Nitrogen gas (N₂) and hydrogen gas (H₂) react to form ammonia gas (NH₃).

   Consider the mixture of N₂ (\(\text{N}_2\)) and H₂ (\(\text{H}_2\)) in a closed container as illustrated below:

   ![Diagram of N₂ and H₂ molecules]

   Assuming the reaction goes to completion, draw a representation of the product mixture. Explain how you arrived at this representation.

5. For the preceding question, which of the following equations best represents the reaction?
6. You know that chemical $A$ reacts with chemical $B$. You react 10.0g $A$ with 10.0g $B$. What information do you need to determine the amount of product that will be produced? Explain.

7. A new grill has a mass of 30.0kg. You put 3.0kg of charcoal in the grill. You burn all the charcoal and the grill has a mass of 30.0kg. What is the mass of the gases given off? Explain.

8. Consider an iron bar on a balance as shown.

As the iron bar rusts, which of the following is true? Explain your answer.

   a. The balance will read less than 75.0g.

   b. The balance will read 75.0g.

   c. The balance will read greater than 75.0g.

   d. The balance will read greater than 75.0g, but if the bar is removed, the rust is scraped off, and the bar replaced, the balance will read 75.0g.

9. You may have noticed that water sometimes drips from the exhaust of a car as it is running. Is this evidence that there is at least a small amount of water originally present in the gasoline? Explain.

Questions 10 and 11 deal with the following situation: You react chemical $A$ with chemical $B$ to make one product. It takes 100g of $A$ to react completely with 20g of $B$. 

\[ \text{6}N_2 + 6H_2 \rightarrow 4NH_3 + 4N_2 \]

\[ N_2 + H_2 \rightarrow NH_3 \]

\[ N + 3H \rightarrow NH_3 \]

\[ N_2 + 3H_2 \rightarrow 2NH_3 \]

\[ 2N_2 + 6H_2 \rightarrow 4NH_3 \]
10. What is the mass of the product?
   a. less than 10g
   b. between 20 and 100g
   c. between 100 and 120g
   d. exactly 120g
   e. more than 120g

11. What is true about the chemical properties of the product?
   a. The properties are more like chemical A.
   b. The properties are more like chemical B.
   c. The properties are an average of those of chemical A and chemical B.
   d. The properties are not necessarily like either chemical A or chemical B.
   e. The properties are more like chemical A or more like chemical B, but more information is needed.

Justify your choice, and for choices you did not pick, explain what is wrong with them.

12. Is there a difference between a homogeneous mixture of hydrogen and oxygen in a 2:1 mole ratio and a sample of water vapor? Explain.

13. Chlorine exists mainly as two isotopes, $^{37}$Cl and $^{35}$Cl. Which is more abundant? How do you know?

14. The average mass of a carbon atom is 12.011. Assuming you could pick up one carbon atom, estimate the chance that you would randomly get one with a mass of 12.011. Support your answer.

15. Can the subscripts in a chemical formula be fractions? Explain. Can the coefficients in a balanced chemical equation be fractions? Explain. Changing the subscripts of chemicals can balance the equations mathematically. Why is this unacceptable?

16. Consider the equation $2A + B \rightarrow A_2B$. If you mix 1.0 mole of $A$ with 1.0 mole of $B$, what amount (moles) of $A_2B$ can be produced?

17. According to the law of conservation of mass, mass cannot be gained or
destroyed in a chemical reaction. Why can't you simply add the masses of two reactants to determine the total mass of product?

18. Which of the following pairs of compounds have the same empirical formula?

   a. acetylene, $C_2H_2$, and benzene, $C_6H_6$
   b. ethane, $C_2H_6$, and butane, $C_4H_{10}$
   c. nitrogen dioxide, $NO_2$, and dinitrogen tetroxide, $N_2O_4$
   d. diphenyl ether, $C_{12}H_{10}O$, and phenol, $C_6H_5OH$

19. Atoms of three different elements are represented by $\text{O}$, $\Box$, and $\triangle$. Which compound is left over when three molecules of $\text{O}\triangle$ and three molecules of $\Box\Box\triangle$ react to form $\text{O}\Box\triangle$ and $\text{O}\Box\Box\triangle$?

20. In chemistry, what is meant by the term “mole”? What is the importance of the mole concept?

21. Which (if any) of the following is (are) true regarding the limiting reactant in a chemical reaction?

   a. The limiting reactant has the lowest coefficient in a balanced equation.
   b. The limiting reactant is the reactant for which you have the fewest number of moles.
   c. The limiting reactant has the lowest ratio of moles available/coefficient in the balanced equation.
   d. The limiting reactant has the lowest ratio of coefficient in the balanced equation/moles available.

   Justify your choice. For those you did not choose, explain why they are incorrect.

22. Consider the equation $3A + B \rightarrow C + D$. You react 4 moles of $A$ with 2 moles of $B$. Which of the following is true?

   a. The limiting reactant is the one with the higher molar mass.
   b. $A$ is the limiting reactant because you need 6 moles of $A$ and have 4 moles.
   c. $B$ is the limiting reactant because you have fewer moles of $B$ than $A$. 
d. B is the limiting reactant because three A molecules react with each B molecule.

e. Neither reactant is limiting.

Justify your choice. For those you did not choose, explain why they are incorrect.

For Review

Questions

Questions with answers below also have full solutions in the Student Solutions Guide, as found on PowerLecture.

23. Reference Section 3.2 to find the atomic masses of $^{12}\text{C}$ and $^{13}\text{C}$, the relative abundance of $^{12}\text{C}$ and $^{13}\text{C}$ in natural carbon, and the average mass (in u) of a carbon atom. If you had a sample of natural carbon containing exactly 10,000 atoms, determine the number of $^{12}\text{C}$ and $^{13}\text{C}$ atoms present. What would be the average mass (in u) and the total mass (in u) of the carbon atoms in this 10,000-atom sample? If you had a sample of natural carbon containing $6.0221 \times 10^{23}$ atoms, determine the number of $^{12}\text{C}$ and $^{13}\text{C}$ atoms present. What would be the average mass (in u) and the total mass (in u) of this $6.0221 \times 10^{23}$ atom sample? Given that $1\text{g} = 6.0221 \times 10^{23}\text{u}$, what is the total mass of 1mole of natural carbon in units of grams?

24. Avogadro’s number, molar mass, and the chemical formula of a compound are three useful conversion factors. What unit conversions can be accomplished using these conversion factors?

25. If you had a mole of U.S. dollar bills and equally distributed the money to all of the people of the world, how rich would every person be? Assume a world population of 7 billion.

26. Describe 1mole of CO$_2$ in as many ways as you can.

27. Which of the following compounds have the same empirical formulas?
28. What is the difference between the molar mass and the empirical formula mass of a compound? When are these masses the same, and when are they different? When different, how is the molar mass related to the empirical formula mass?

29. How is the mass percent of elements in a compound different for a 1.0-g sample versus a 100-g sample versus a 1-mole sample of the compound?

30. A balanced chemical equation contains a large amount of information. What information is given in a balanced equation?

31. The reaction of an element X with element Y is represented in the following diagram. Which of the equations best describes this reaction?

- a. \(3X + 8Y \rightarrow X_3Y_8\)
- b. \(3X + 6Y \rightarrow X_3Y_6\)
- c. \(X + 2Y \rightarrow XY_2\)
- d. \(3X + 8Y \rightarrow 3XY_2 + 2Y\)
32. Hydrogen gas and oxygen gas react to form water, and this reaction can be depicted as follows:

\[ \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \]

Explain why this equation is not balanced, and draw a picture of the balanced equation.

33. What is the theoretical yield for a reaction, and how does this quantity depend on the limiting reactant?

34. What does it mean to say a reactant is present “in excess” in a process? Can the limiting reactant be present in excess? Does the presence of an excess of a reactant affect the mass of products expected for a reaction?

35. Consider the following generic reaction:

\[ \text{A}_2\text{B}_2 + 2\text{C} \rightarrow 2\text{CB} + 2\text{A} \]

What steps and information are necessary to perform the following determinations assuming that 1.00 \( \times \) 10\(^4\) molecules of \( \text{A}_2\text{B}_2 \) are reacted with excess \( \text{C} \)?

- a. mass of \( \text{CB} \) produced
- b. atoms of \( \text{A} \) produced
- c. moles of \( \text{C} \) reacted
- d. percent yield of \( \text{CB} \)

36. Consider the following generic reaction:

\[ \text{Y}_2 + 2\text{XY} \rightarrow 2\text{XY}_2 \]

In a limiting reactant problem, a certain quantity of each reactant is given and you are usually asked to calculate the mass of product formed. If 10.0g of \( \text{Y}_2 \) is reacted with 10.0g of \( \text{XY} \), outline two methods you could use to determine which reactant is limiting (runs out first) and thus determines the mass of product formed.
37. An element consists of 1.40% of an isotope with mass 203.973u, 24.10% of an isotope with mass 205.9745u, 22.10% of an isotope with mass 206.9759u, and 52.40% of an isotope with mass 207.9766u. Calculate the average atomic mass, and identify the element.

38. An element “X” has five major isotopes, which are listed below along with their abundances. What is the element?

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Percent Natural Abundance</th>
<th>Mass (u)</th>
</tr>
</thead>
<tbody>
<tr>
<td>46 X</td>
<td>8.00%</td>
<td>45.95232</td>
</tr>
<tr>
<td>47 X</td>
<td>7.30%</td>
<td>46.951764</td>
</tr>
<tr>
<td>48 X</td>
<td>73.80%</td>
<td>47.947947</td>
</tr>
<tr>
<td>49 X</td>
<td>5.50%</td>
<td>48.947841</td>
</tr>
<tr>
<td>50 X</td>
<td>5.40%</td>
<td>49.944792</td>
</tr>
</tbody>
</table>

39. The element rhenium (Re) has two naturally occurring isotopes, $^{185}$Re and $^{187}$Re, with an average atomic mass of 186.207u. Rhenium is 62.60%$^{187}$Re, and the atomic mass of $^{187}$Re is 186.956u. Calculate the mass of $^{185}$Re.

40. Assume silicon has three major isotopes in nature as shown in the table below. Fill in the missing information.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (u)</th>
<th>Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>28 Si</td>
<td>27.98</td>
<td>$\text{_______}$</td>
</tr>
</tbody>
</table>
41. The element europium exists in nature as two isotopes: $^{151}\text{Eu}$ has a mass of 150.9196\,\text{u} and $^{153}\text{Eu}$ has a mass of 152.9209\,\text{u}. The average atomic mass of europium is 151.96\,\text{u}. Calculate the relative abundance of the two europium isotopes.

42. The element silver (Ag) has two naturally occurring isotopes: $^{109}\text{Ag}$ and $^{107}\text{Ag}$ with a mass of 106.905\,\text{u}. Silver consists of 51.82\%$^{107}\text{Ag}$ and has an average atomic mass of 107.868\,\text{u}. Calculate the mass of $^{109}\text{Ag}$.

43. The mass spectrum of bromine (Br$_2$) consists of three peaks with the following characteristics:

<table>
<thead>
<tr>
<th>Mass (u)</th>
<th>Relative Size</th>
</tr>
</thead>
<tbody>
<tr>
<td>157.84</td>
<td>0.2534</td>
</tr>
<tr>
<td>159.84</td>
<td>0.5000</td>
</tr>
<tr>
<td>161.84</td>
<td>0.2466</td>
</tr>
</tbody>
</table>

How do you interpret these data?

44. The stable isotopes of iron are $^{54}\text{Fe}$, $^{56}\text{Fe}$, $^{57}\text{Fe}$, and $^{58}\text{Fe}$. The mass spectrum of iron looks like the following:
Use the data on the mass spectrum to estimate the average atomic mass of iron, and compare it to the value given in the table inside the front cover of this book.

Chapter 3: Stoichiometry Exercises: Moles and Molar Masses

For Review

Exercises: Moles and Molar Masses

45. Calculate the mass of 500. atoms of iron (Fe).

46. What number of Fe atoms and what amount (moles) of Fe atoms are in 500.0 g of iron?

47. Diamond is a natural form of pure carbon. What number of atoms of carbon are in a 1.00-carat diamond (1.00 carat = 0.200 g)?

48. A diamond contains $5.0 \times 10^{21}$ atoms of carbon. What amount (moles) of carbon and what mass (grams) of carbon are in this diamond?

49. Aluminum metal is produced by passing an electric current through a solution of aluminum oxide ($\text{Al}_2\text{O}_3$) dissolved in molten cryolite ($\text{Na}_3\text{AlF}_6$). Calculate the molar masses of $\text{Al}_2\text{O}_3$ and $\text{Na}_3\text{AlF}_6$.

50. The Freons are a class of compounds containing carbon, chlorine, and fluorine. While they have many valuable uses, they have been shown to be responsible for depletion of the ozone in the upper atmosphere. In 1991, two replacement compounds for Freons went into production: HFC-134a($\text{CH}_2\text{FCF}_3$) and HCFC-124($\text{CHClF}_2\text{CF}_3$). Calculate the molar masses of these two compounds.

51. Calculate the molar mass of the following substances.

a. \[ \text{H} \quad \text{N} \]

b. \[ \text{H} \quad \text{N} \]

c. $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$
52. Calculate the molar mass of the following substances.

a. 

![Diagram of a compound with atoms O and P]

b. \( \text{Ca}_3(\text{PO}_4)_2 \)

c. \( \text{Na}_2\text{HPO}_4 \)

53. What amount (moles) of compound is present in 1.00g of each of the compounds in Exercise 51?

54. What amount (moles) of compound is present in 1.00g of each of the compounds in Exercise 52?

55. What mass of compound is present in 5.00 moles of each of the compounds in Exercise 51?

56. What mass of compound is present in 5.00 moles of each of the compounds in Exercise 52?

57. What mass of nitrogen is present in 5.00 moles of each of the compounds in Exercise 51?

58. What mass of phosphorus is present in 5.00 moles of each of the compounds in Exercise 52?

59. What number of molecules (or formula units) are present in 1.00g of each of the compounds in Exercise 51?

60. What number of molecules (or formula units) are present in 1.00g of each of the compounds in Exercise 52?

61. What number of atoms of nitrogen are present in 1.00g of each of the compounds in Exercise 51?

62. What number of atoms of phosphorus are present in 1.00g of each of the compounds in Exercise 52?

63. Freon-12(\( \text{CCl}_2\text{F}_2 \)) is used as a refrigerant in air conditioners and as a propellant in aerosol cans. Calculate the number of molecules of Freon-12 in 5.56 mg of Freon-12. What is the mass of chlorine in 5.56 mg of Freon-12?

64. Bauxite, the principal ore used in the production of aluminum, has a molecular formula of \( \text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O} \). The \( \cdot \text{H}_2\text{O} \) in the formula are called
waters of hydration. Each formula unit of the compound contains two water molecules.

a. What is the molar mass of bauxite?

b. What is the mass of aluminum in 0.58 mole of bauxite?

c. How many atoms of aluminum are in 0.58 mole of bauxite?

d. What is the mass of $2.1 \times 10^{24}$ formula units of bauxite?

65. What amount (moles) is represented by each of these samples?

a. 150.0g Fe₂O₃

b. 10.0mg NO₂

c. $1.5 \times 10^{16}$ molecules of BF₃

66. What amount (moles) is represented by each of these samples?

a. 20.0mg caffeine, C₈H₁₀N₄O₂

b. $2.72 \times 10^{21}$ molecules of ethanol, C₂H₅OH

c. 1.50g of dry ice, CO₂

67. What number of atoms of nitrogen are present in 5.00g of each of the following?

a. glycine, C₂H₅O₂N

b. magnesium nitride

c. calcium nitrate

d. dinitrogen tetroxide

68. Complete the following table.

<table>
<thead>
<tr>
<th>Mass of Sample</th>
<th>Moles of Sample</th>
<th>Molecules in Sample</th>
<th>Total Atoms in Sample</th>
</tr>
</thead>
<tbody>
<tr>
<td>4.24g C₆H₆</td>
<td>_______</td>
<td>_______</td>
<td>_______</td>
</tr>
<tr>
<td>_______</td>
<td>0.224mol H₂O</td>
<td>_______</td>
<td>_______</td>
</tr>
</tbody>
</table>
69. Ascorbic acid, or vitamin C(C₆H₈O₆), is an essential vitamin. It cannot be stored by the body and must be present in the diet. What is the molar mass of ascorbic acid? Vitamin C tablets are taken as a dietary supplement. If a typical tablet contains 500.0 mg vitamin C, what amount (moles) and what number of molecules of vitamin C does it contain?

70. The molecular formula of acetylsalicylic acid (aspirin), one of the most commonly used pain relievers, is C₉H₈O₄.
   a. Calculate the molar mass of aspirin.
   b. A typical aspirin tablet contains 500.0 mg C₉H₈O₄. What amount (moles) of C₉H₈O₄ molecules and what number of molecules of acetylsalicylic acid are in a 500.0 mg tablet?

71. Chloral hydrate (C₂H₃Cl₃O₂) is a drug formerly used as a sedative and hypnotic. It is the compound used to make “Mickey Finns” in detective stories.
   a. Calculate the molar mass of chloral hydrate.
   b. What amount (moles) of C₂H₃Cl₃O₂ molecules are in 500.0 g chloral hydrate?
   c. What is the mass in grams of 2.0 × 10⁻² mol of chloral hydrate?
   d. What number of chlorine atoms are in 5.0 g chloral hydrate?
   e. What mass of chloral hydrate would contain 1.0 g Cl?
   f. What is the mass of exactly 500 molecules of chloral hydrate?

72. Dimethylnitrosamine, \((\text{CH}_3)_2\text{N}_2\text{O}\), is a carcinogenic (cancer-causing) substance that may be formed in foods, beverages, or gastric juices from the reaction of nitrite ion (used as a food preservative) with other substances.
73. Calculate the percent composition by mass of the following compounds that are important starting materials for synthetic polymers:

a. C₃H₄O₂ (acrylic acid, from which acrylic plastics are made)

b. C₄H₆O₂ (methyl acrylate, from which Plexiglas is made)

c. C₃H₃N (acrylonitrile, from which Orlon is made)

74. In 1987 the first substance to act as a superconductor at a temperature above that of liquid nitrogen (77K) was discovered. The approximate formula of this substance is YBa₂Cu₃O₇. Calculate the percent composition by mass of this material.

75. The percent by mass of nitrogen for a compound is found to be 46.7%. Which of the following could be this species?

76. Arrange the following substances in order of increasing mass percent of carbon.
Fungal laccase, a blue protein found in wood-rotting fungi, is 0.390% Cu by mass. If a fungal laccase molecule contains four copper atoms, what is the molar mass of fungal laccase?

Hemoglobin is the protein that transports oxygen in mammals. Hemoglobin is 0.347% Fe by mass, and each hemoglobin molecule contains four iron atoms. Calculate the molar mass of hemoglobin.

Express the composition of each of the following compounds as the mass percents of its elements.

a. formaldehyde, CH₂O

b. glucose, C₆H₁₂O₆

c. acetic acid, CH₃COOH

Considering your answer to Exercise 79, which type of formula, empirical or molecular, can be obtained from elemental analysis that gives percent composition?

Give the empirical formula for each of the compounds represented below.

a. [Diagram]

b. [Diagram]
82. Determine the molecular formulas to which the following empirical formulas and molar masses pertain.

a. SNH(188.35g/mol)

b. NPCl₂ (347.64g/mol)

c. CoC₄O₄ (341.94g/mol)

d. SN (184.32g/mol)

83. A compound that contains only carbon, hydrogen, and oxygen is 48.64% C and 8.16% H by mass. What is the empirical formula of this substance?

84. The most common form of nylon (nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula for nylon-6.

85. There are two binary compounds of mercury and oxygen. Heating either of them results in the decomposition of the compound, with oxygen gas escaping into the atmosphere while leaving a residue of pure mercury. Heating 0.6498g of one of the compounds leaves a residue of 0.6018g. Heating 0.4172g of the other compound results in a mass loss of 0.016g. Determine the empirical formula of each compound.

86. A sample of urea contains 1.121g N, 0.161g H, 0.480g C, and 0.640g O. What is the empirical formula of urea?

87. A compound containing only sulfur and nitrogen is 69.6% S by mass; the molar mass is 184 g/mol. What are the empirical and molecular formulas of the compound?
88. Determine the molecular formula of a compound that contains 26.7% P, 12.1% N, and 61.2% Cl, and has a molar mass of 580 g/mol.

89. A compound contains 47.08% carbon, 6.59% hydrogen, and 46.33% chlorine by mass; the molar mass of the compound is 153 g/mol. What are the empirical and molecular formulas of the compound?

90. Maleic acid is an organic compound composed of 41.39% C, 3.47% H, and the rest oxygen. If 0.129 mole of maleic acid has a mass of 15.0 g, what are the empirical and molecular formulas of maleic acid?

91. One of the components that make up common table sugar is fructose, a compound that contains only carbon, hydrogen, and oxygen. Complete combustion of 1.50 g of fructose produced 2.20 g of carbon dioxide and 0.900 g of water. What is the empirical formula of fructose?

92. A compound contains only C, H, and N. Combustion of 35.0 mg of the compound produces 33.5 mg CO₂ and 41.1 mg H₂O. What is the empirical formula of the compound?

93. Cumene is a compound containing only carbon and hydrogen that is used in the production of acetone and phenol in the chemical industry. Combustion of 47.6 mg cumene produces some CO₂ and 42.8 mg water. The molar mass of cumene is between 115 and 125 g/mol. Determine the empirical and molecular formulas.

94. A compound contains only carbon, hydrogen, and oxygen. Combustion of 10.68 mg of the compound yields 16.01 mg CO₂ and 4.37 mg H₂O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of the compound?

Chapter 3: Stoichiometry Exercises: Balancing Chemical Equations
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For Review

Exercises: Balancing Chemical Equations

95. Give the balanced equation for each of the following chemical reactions:

   a. Glucose (C₆H₁₂O₆) reacts with oxygen gas to produce gaseous carbon dioxide and water vapor.
b. Solid iron(III) sulfide reacts with gaseous hydrogen chloride to form solid iron(III) chloride and hydrogen sulfide gas.

c. Carbon disulfide liquid reacts with ammonia gas to produce hydrogen sulfide gas and solid ammonium thiocyanate (NH₄SCN).

96. Give the balanced equation for each of the following.

a. The combustion of ethanol (C₂H₅OH) forms carbon dioxide and water vapor. A combustion reaction refers to a reaction of a substance with oxygen gas.

b. Aqueous solutions of lead(II) nitrate and sodium phosphate are mixed, resulting in the precipitate formation of lead(II) phosphate with aqueous sodium nitrate as the other product.

c. Solid zinc reacts with aqueous HCl to form aqueous zinc chloride and hydrogen gas.

d. Aqueous strontium hydroxide reacts with aqueous hydro-bromic acid to produce water and aqueous strontium bromide.

97. A common demonstration in chemistry courses involves adding a tiny speck of manganese(IV) oxide to a concentrated hydrogen peroxide (H₂O₂) solution. Hydrogen peroxide decomposes quite spectacularly under these conditions to produce oxygen gas and steam (water vapor). Manganese(IV) oxide is a catalyst for the decomposition of hydrogen peroxide and is not consumed in the reaction. Write the balanced equation for the decomposition reaction of hydrogen peroxide.

98. Iron oxide ores, commonly a mixture of FeO and Fe₂O₃, are given the general formula Fe₃O₄. They yield elemental iron when heated to a very high temperature with either carbon monoxide or elemental hydrogen. Balance the following equations for these processes:

\[
\text{Fe}_3\text{O}_4(s) + \text{H}_2(g) \rightarrow \text{Fe}(s) + \text{H}_2\text{O}(g)
\]

\[
\text{Fe}_3\text{O}_4(s) + \text{CO}(g) \rightarrow \text{Fe}(s) + \text{CO}_2(g)
\]

99. Balance the following equations:

a. Ca(OH)₂(aq) + H₃PO₄(aq) → H₂O(l) + Ca₃(PO₄)₂(s)

b. Al(OH)₃(s) + HCl(aq) → AlCl₃(aq) + H₂O(l)

c. AgNO₃(aq) + H₂SO₄(aq) → Ag₂SO₄(s) + HNO₃(aq)

100. Balance each of the following chemical equations.
Balance the following equations representing combustion reactions:

a. \( \text{KO}_2(s) + \text{H}_2\text{O}(l) \rightarrow \text{KOH}(aq) + \text{O}_2(g) + \text{H}_2\text{O}_2(aq) \)

b. \( \text{Fe}_2\text{O}_3(s) + \text{HNO}_3(aq) \rightarrow \text{Fe(NO}_3)_3(aq) + \text{H}_2\text{O}(l) \)

c. \( \text{NH}_3(g) + \text{O}_2(g) \rightarrow \text{NO}(g) + \text{H}_2\text{O}(g) \)

d. \( \text{PCl}_5(l) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{PO}_4(aq) + \text{HCl}(g) \)

e. \( \text{CaO}(s) + \text{C}(s) \rightarrow \text{CaC}_2(s) + \text{CO}_2(g) \)

f. \( \text{MoS}_2(s) + \text{O}_2(g) \rightarrow \text{MoO}_3(s) + \text{SO}_2(g) \)

g. \( \text{FeCO}_3(s) + \text{H}_2\text{CO}_3(aq) \rightarrow \text{Fe(HCO}_3)_2(aq) \)

101. Balance the following equations representing combustion reactions:

a. \[ \begin{array}{c}
\text{H}_2\text{C}_2\text{O}_4(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)
\end{array} \]

d. \[ \begin{array}{c}
\text{Fe}(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)
\end{array} \]

e. \[ \begin{array}{c}
\text{FeO}(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)
\end{array} \]

102. Balance the following equations:

a. \( \text{Cr}(s) + \text{S}_8(s) \rightarrow \text{Cr}_2\text{S}_3(s) \)

b. \( \text{NaHCO}_3(s) \xrightarrow{\text{Heat}} \text{Na}_2\text{CO}_3(s) + \text{CO}_2(g) + \text{H}_2\text{O}(g) \)

c. \( \text{KClO}_3(s) \xrightarrow{\text{Heat}} \text{KCl}(s) + \text{O}_2(g) \)

d. \( \text{Eu}(s) + \text{HF}(g) \rightarrow \text{EuF}_3(s) + \text{H}_2(g) \)

103. Silicon is produced for the chemical and electronics industries by the following reactions. Give the balanced equation for each reaction.

a. \( \text{SiO}_2(s) + \text{C}(s) \xrightarrow{\text{Electric arc furnace}} \text{Si}(s) + \text{CO}(g) \)

b. Liquid silicon tetrachloride is reacted with very pure solid magnesium, producing solid silicon and solid magnesium chloride.

c. \( \text{Na}_2\text{SiF}_6(s) + \text{Na}(s) \rightarrow \text{Si}(s) + \text{NaF}(s) \)
104. Glass is a mixture of several compounds, but a major constituent of most glass is calcium silicate, \( \text{CaSiO}_3 \). Glass can be etched by treatment with hydrofluoric acid; \( \text{HF} \) attacks the calcium silicate of the glass, producing gaseous and water-soluble products (which can be removed by washing the glass). For example, the volumetric glassware in chemistry laboratories is often graduated by using this process. Balance the following equation for the reaction of hydrofluoric acid with calcium silicate.

\[
\text{CaSiO}_3(s) + \text{HF}(aq) \rightarrow \text{CaF}_2(aq) + \text{SiF}_4(g) + \text{H}_2\text{O}(l)
\]

Chapter 3: Stoichiometry Exercises: Reaction Stoichiometry
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**For Review**

**Exercises: Reaction Stoichiometry**

105. Over the years, the thermite reaction has been used for welding railroad rails, in incendiary bombs, and to ignite solid-fuel rocket motors. The reaction is

\[
\text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow 2\text{Fe}(l) + \text{Al}_2\text{O}_3(s)
\]

What masses of iron(III) oxide and aluminum must be used to produce 15.0g iron? What is the maximum mass of aluminum oxide that could be produced?

106. The reaction between potassium chlorate and red phosphorus takes place when you strike a match on a matchbox. If you were to react 52.9g of potassium chlorate (\( \text{KClO}_3 \)) with excess red phosphorus, what mass of tetraphosphorus decaoxide (\( \text{P}_4\text{O}_{10} \)) could be produced?

\[
\text{KClO}_3(s) + \text{P}_4(s) \rightarrow \text{P}_4\text{O}_{10}(s) + \text{KCl}(s) \quad \text{(unbalanced)}
\]

107. The reusable booster rockets of the U.S. space shuttle employ a mixture of aluminum and ammonium perchlorate for fuel. A possible equation for this reaction is

\[
3\text{Al}(s) + 3\text{NH}_4\text{ClO}_4(s) \rightarrow \text{Al}_2\text{O}_3(s) + \text{AlCl}_3(s) + 3\text{NO}(g) + 6\text{H}_2\text{O}(g)
\]

What mass of \( \text{NH}_4\text{ClO}_4 \) should be used in the fuel mixture for every kilogram
One of relatively few reactions that takes place directly between two solids at room temperature is

$$\text{Ba(OH)}_2 \cdot 8\text{H}_2\text{O}(s) + \text{NH}_4\text{SCN}(s) \rightarrow \text{Ba(SCN)}_2(s) + \text{H}_2\text{O}(l) + \text{NH}_3(g)$$

In this equation, the $\cdot 8\text{H}_2\text{O}$ in $\text{Ba(OH)}_2 \cdot 8\text{H}_2\text{O}$ indicates the presence of eight water molecules. This compound is called barium hydroxide octahydrate.

a. Balance the equation.

b. What mass of ammonium thiocyanate (NH$_4$SCN) must be used if it is to react completely with 6.5g barium hydroxide octahydrate?

109. Elixirs such as Alka-Seltzer use the reaction of sodium bicarbonate with citric acid in aqueous solution to produce a fizz:

$$3\text{NaHCO}_3(aq) + \text{C}_6\text{H}_8\text{O}_7(aq) \rightarrow 3\text{CO}_2(g) + 3\text{H}_2\text{O}(l) + \text{Na}_3\text{C}_6\text{H}_5\text{O}_7(aq)$$

a. What mass of C$_6$H$_8$O$_7$ should be used for every $1.0 \times 10^2\text{mgNaHCO}_3$?

b. What mass of CO$_2(g)$ could be produced from such a mixture?

110. Aspirin (C$_9$H$_8$O$_4$) is synthesized by reacting salicylic acid (C$_7$H$_6$O$_3$) with acetic anhydride (C$_4$H$_6$O$_3$). The balanced equation is

$$\text{C}_7\text{H}_6\text{O}_3 + \text{C}_4\text{H}_6\text{O}_3 \rightarrow \text{C}_9\text{H}_8\text{O}_4 + \text{HC}_2\text{H}_5\text{O}_2$$

a. What mass of acetic anhydride is needed to completely consume $1.00 \times 10^2\text{g}$ salicylic acid?

b. What is the maximum mass of aspirin (the theoretical yield) that could be produced in this reaction?

111. Bacterial digestion is an economical method of sewage treatment. The reaction

$$5\text{CO}_2(g) + 55\text{NH}_4^+(aq) + 76\text{O}_2(g) \xrightarrow{\text{bacteria}} \text{C}_5\text{H}_7\text{O}_2\text{N}(s) + 54\text{NO}_2^-(aq) + 52\text{H}_2\text{O}(l) + 109\text{H}^+(aq)$$

is an intermediate step in the conversion of the nitrogen in organic compounds into nitrate ions. What mass of bacterial tissue is produced in a treatment plant for every $1.0 \times 10^4\text{kg}$ of wastewater containing 3.0%NH$_4^+$ ions by mass? Assume that 95% of the ammonium ions are consumed by the
112. Phosphorus can be prepared from calcium phosphate by the following reaction:

\[
2\text{Ca}_3(\text{PO}_4)_2(s) + 6\text{SiO}_2(s) + 10\text{C}(s) \rightarrow 6\text{CaSiO}_3(s) + \text{P}_4(s) + 10\text{CO}(g)
\]

Phosphorite is a mineral that contains \(\text{Ca}_3(\text{PO}_4)_2\) plus other non-phosphorus-containing compounds. What is the maximum amount of \(\text{P}_4\) that can be produced from 1.0 kg of phosphorite if the phorphorite sample is 75% \(\text{Ca}_3(\text{PO}_4)_2\) by mass? Assume an excess of the other reactants.

113. Coke is an impure form of carbon that is often used in the industrial production of metals from their oxides. If a sample of coke is 95% carbon by mass, determine the mass of coke needed to react completely with 1.0 ton of copper(II) oxide.

\[
2\text{CuO}(s) + \text{C}(s) \rightarrow 2\text{Cu}(s) + \text{CO}_2(g)
\]

114. The space shuttle environmental control system handles excess \(\text{CO}_2\) (which the astronauts breathe out; it is 4.0% by mass of exhaled air) by reacting it with lithium hydroxide, \(\text{LiOH}\), pellets to form lithium carbonate, \(\text{Li}_2\text{CO}_3\), and water. If there are seven astronauts on board the shuttle, and each exhales 20 L of air per minute, how long could clean air be generated if there were 25,000 g of \(\text{LiOH}\) pellets available for each shuttle mission? Assume the density of air is 0.0010 g/mL.

### For Review

**Exercises: Limiting Reactants and Percent Yield**

115. Consider the reaction between \(\text{NO}(g)\) and \(\text{O}_2(g)\) represented below.
What is the balanced equation for this reaction, and what is the limiting reactant?

116. Consider the following reaction:

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

If a container were to have 10 molecules of $\text{O}_2$ and 10 molecules of $\text{NH}_3$ initially, how many total molecules (reactants plus products) would be present in the container after this reaction goes to completion?

117. Ammonia is produced from the reaction of nitrogen and hydrogen according to the following balanced equation:

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

a. What is the maximum mass of ammonia that can be produced from a mixture of $1.00 \times 10^3 \text{gN}_2$ and $5.00 \times 10^2 \text{gH}_2$?

b. What mass of which starting material would remain unreacted?

118. Consider the following unbalanced equation:

$$\text{Ca}_3(\text{PO}_4)_2(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CaSO}_4(s) + \text{H}_3\text{PO}_4(aq)$$

What masses of calcium sulfate and phosphoric acid can be produced from the reaction of 1.0kg calcium phosphate with 1.0kg concentrated sulfuric acid (98%$\text{H}_2\text{SO}_4$ by mass)?

119. Hydrogen peroxide is used as a cleansing agent in the treatment of cuts and abrasions for several reasons. It is an oxidizing agent that can directly kill many microorganisms; it decomposes on contact with blood, releasing elemental oxygen gas (which inhibits the growth of anaerobic microorganisms); and it foams on contact with blood, which provides a cleansing action. In the laboratory, small quantities of hydrogen peroxide can be prepared by the action of an acid on an alkaline earth metal peroxide, such as barium peroxide:

$$\text{BaO}_2(s) + 2\text{HCl}(aq) \rightarrow \text{H}_2\text{O}_2(aq) + \text{BaCl}_2(aq)$$
What mass of hydrogen peroxide should result when 1.50 g barium peroxide is treated with 25.0 mL hydrochloric acid solution containing 0.0272 g HCl per mL? What mass of which reagent is left unreacted?

120. Silver sulfadiazine burn-treating cream creates a barrier against bacterial invasion and releases antimicrobial agents directly into the wound. If 25.0 g Ag₂O is reacted with 50.0 g C₁₀H₁₀N₄SO₂, what mass of silver sulfadiazine, AgC₁₀H₉N₄SO₂, can be produced, assuming 100% yield?

\[
\text{Ag}_2\text{O}(s) + 2\text{C}_{10}\text{H}_{10}\text{N}_4\text{SO}_2(s) \rightarrow 2\text{AgC}_{10}\text{H}_9\text{N}_4\text{SO}_2(s) + \text{H}_2\text{O}(l)
\]

121. Hydrogen cyanide is produced industrially from the reaction of gaseous ammonia, oxygen, and methane:

\[
2\text{NH}_3(g) + 3\text{O}_2(g) + 2\text{CH}_4(g) \rightarrow 2\text{HCN}(g) + 6\text{H}_2\text{O}(g)
\]

If \(5.00 \times 10^3\) kg each of NH₃, O₂, and CH₄ are reacted, what mass of HCN and H₂O will be produced, assuming 100% yield?

122. Acrylonitrile (C₃H₃N) is the starting material for many synthetic carpets and fabrics. It is produced by the following reaction.

\[
2\text{C}_3\text{H}_6(g) + 2\text{NH}_3(g) + 3\text{O}_2(g) \rightarrow 2\text{C}_3\text{H}_3\text{N}(g) + 6\text{H}_2\text{O}(g)
\]

If 15.0 g C₃H₆, 10.0 g O₂, and 5.00 g NH₃ are reacted, what mass of acrylonitrile can be produced, assuming 100% yield?

123. The reaction of ethane gas (C₂H₆) with chlorine gas produces C₂H₅Cl as its main product (along with HCl). In addition, the reaction invariably produces a variety of other minor products, including C₂H₄Cl₂, C₂H₃Cl₃, and others. Naturally, the production of these minor products reduces the yield of the main product. Calculate the percent yield of C₂H₅Cl if the reaction of 300 g of ethane with 650 g of chlorine produced 490 g of C₂H₅Cl.

124. DDT, an insecticide harmful to fish, birds, and humans, is produced by the following reaction:

\[
2\text{C}_6\text{H}_₅\text{Cl} + \text{C}_₂\text{HOCl}_₃ \rightarrow \text{C}_{₁₄}\text{H}_₉\text{Cl}_₅ + \text{H}_₂\text{O}
\]

In a government lab, 1142 g of chlorobenzene is reacted with 485 g of chloral.

a. What mass of DDT is formed, assuming 100% yield?

b. Which reactant is limiting? Which is in excess?

c. What mass of the excess reactant is left over?
d. If the actual yield of DDT is 200.0 g, what is the percent yield?

125. Bornite \((\text{Cu}_3\text{Fe}_3\text{S}_3)\) is a copper ore used in the production of copper. When heated, the following reaction occurs:

\[
2\text{Cu}_3\text{Fe}_3\text{S}_3(s) + 7\text{O}_2(g) \rightarrow 6\text{Cu}(s) + 2\text{FeO}(s) + 6\text{SO}_2(g)
\]

If 2.50 metric tons of bornite is reacted with excess \(\text{O}_2\) and the process has an 86.3\% yield of copper, what mass of copper is produced?

126. Consider the following unbalanced reaction:

\[
P_4(s) + \text{F}_2(g) \rightarrow \text{PF}_3(g)
\]

What mass of \(\text{F}_2\) is needed to produce 120. g of \(\text{PF}_3\) if the reaction has a 78.1\% yield?

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Chapter 3: Stoichiometry Additional Exercises

For Review

Additional Exercises

127. In using a mass spectrometer, a chemist sees a peak at a mass of 30.0106. Of the choices \(^{12}\text{C}\)\(^2\text{H}_6\), \(^{12}\text{C}\)\(^1\text{H}_2\)\(^{16}\text{O}\), and \(^{14}\text{N}\)\(^{16}\text{O}\), which is responsible for this peak? Pertinent masses are \(^1\text{H}\), 1.007825; \(^{16}\text{O}\), 15.994915; and \(^{14}\text{N}\), 14.003074.

128. Boron consists of two isotopes, \(^{10}\text{B}\) and \(^{11}\text{B}\). Chlorine also has two isotopes, \(^{35}\text{Cl}\) and \(^{37}\text{Cl}\). Consider the mass spectrum of \(\text{BCl}_3\). How many peaks would be present, and what approximate mass would each peak correspond to in the \(\text{BCl}_3\) mass spectrum?

129. A given sample of a xenon fluoride compound contains molecules of the type \(\text{XeF}_n\), where \(n\) is some whole number. Given that \(9.03 \times 10^{20}\) molecules of \(\text{XeF}_n\) weigh 0.368 g, determine the value for \(n\) in the formula.

130. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (table sugar) when dissolved in water. It is marketed as NutraSweet. The molecular formula of aspartame is \(\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5\).

a. Calculate the molar mass of aspartame.
b. What amount (moles) of molecules are present in 10.0g aspartame?

c. Calculate the mass in grams of 1.56 mole of aspartame.

d. What number of molecules are in 5.0mg aspartame?

e. What number of atoms of nitrogen are in 1.2g aspartame?

f. What is the mass in grams of 1.0 \times 10^9 \text{ molecules of aspartame}?

g. What is the mass in grams of one molecule of aspartame?

131. Anabolic steroids are performance enhancement drugs whose use has been banned from most major sporting activities. One anabolic steroid is fluoxymesterone (C_{20}H_{29}FO_3). Calculate the percent composition by mass of fluoxymesterone.

132. Many cereals are made with high moisture content so that the cereal can be formed into various shapes before it is dried. A cereal product containing 58\% H_2O by mass is produced at the rate of 1000. kg/h. What mass of water must be evaporated per hour if the final product contains only 20\% water?

133. The compound adrenaline contains 56.79\% C, 6.56\% H, 28.37\% O, and 8.28\% N by mass. What is the empirical formula for adrenaline?

134. Adipic acid is an organic compound composed of 49.31\% C, 43.79\% O, and the rest hydrogen. If the molar mass of adipic acid is 146.1g/mol, what are the empirical and molecular formulas for adipic acid?

135. Vitamin B_{12}, cyanocobalamin, is essential for human nutrition. It is concentrated in animal tissue but not in higher plants. Although nutritional requirements for the vitamin are quite low, people who abstain completely from animal products may develop a deficiency anemia. Cyanocobalamin is the form used in vitamin supplements. It contains 4.34\% cobalt by mass. Calculate the molar mass of cyanocobalamin, assuming that there is one atom of cobalt in every molecule of cyanocobalamin.

136. Some bismuth tablets, a medication used to treat upset stomachs, contain 262mg of bismuth subsalicylate, C_7H_5BiO_6, per tablet. Assuming two tablets are digested, calculate the mass of bismuth consumed.

137. The empirical formula of styrene is CH; the molar mass of styrene is 104.14 g/mol. What number of H atoms are present in a 2.00-g sample of styrene?

138. Terephthalic acid is an important chemical used in the manufacture of polyesters and plasticizers. It contains only C, H, and O. Combustion of
19.81 mg terephthalic acid produces 41.98 mg CO₂ and 6.45 mg H₂O. If 0.250 mole of terephthalic acid has a mass of 41.5 g, determine the molecular formula for terephthalic acid.

139. A sample of a hydrocarbon (a compound consisting of only carbon and hydrogen) contains 2.59 \times 10^{23} \text{ atoms of hydrogen and is } 17.3\% \text{ hydrogen by mass. If the molar mass of the hydrocarbon is between 55 and 65 g/mol, what amount (moles) of compound is present, and what is the mass of the sample?}

140. A binary compound between an unknown element E and hydrogen contains 91.27\% E and 8.73\% H by mass. If the formula of the compound is E₃H₈, calculate the atomic mass of E.

141. A 0.755-g sample of hydrated copper(II) sulfate

\[ \text{CuSO}_4 \cdot x\text{H}_2\text{O} \]

was heated carefully until it had changed completely to anhydrous copper(II) sulfate (CuSO₄) with a mass of 0.483 g. Determine the value of \( x \). [This number is called the number of waters of hydration of copper(II) sulfate. It specifies the number of water molecules per formula unit of CuSO₄ in the hydrated crystal.]

142. ABS plastic is a tough, hard plastic used in applications requiring shock resistance. The polymer consists of three monomer units: acrylonitrile (C₃H₃N), butadiene (C₄H₆), and styrene (C₈H₈).

a. A sample of ABS plastic contains 8.80\% N by mass. It took 0.605 g of Br₂ to react completely with a 1.20-g sample of ABS plastic. Bromine reacts 1:1 (by moles) with the butadiene molecules in the polymer and nothing else. What is the percent by mass of acrylonitrile and butadiene in this polymer?

b. What are the relative numbers of each of the monomer units in this polymer?

143. A sample of LSD (d-lysergic acid diethylamide, C₂₄H₃₉N₃O) is added to some table salt (sodium chloride) to form a mixture. Given that a 1.00-g sample of the mixture undergoes combustion to produce 1.20 g of CO₂, what is the mass percent of LSD in the mixture?

144. Methane (CH₄) is the main component of marsh gas. Heating methane in the presence of sulfur produces carbon disulfide and hydrogen sulfide as the only products.

a. Write the balanced chemical equation for the reaction of methane and
b. Calculate the theoretical yield of carbon disulfide when 120.0 g of methane is reacted with an equal mass of sulfur.

145. A potential fuel for rockets is a combination of \( B_5H_9 \) and \( O_2 \). The two react according to the following balanced equation:

\[
2B_5H_9(l) + 12O_2(g) \rightarrow 5B_2O_3(s) + 9H_2O(g)
\]

If one tank in a rocket holds 126.0 g \( B_5H_9 \) and another tank holds 192.0 g \( O_2 \), what mass of water can be produced when the entire contents of each tank react together?

146. A 0.4230 g sample of impure sodium nitrate was heated, converting all the sodium nitrate to 0.2864 g of sodium nitrite and oxygen gas. Determine the percent of sodium nitrate in the original sample.

147. An iron ore sample contains \( Fe_2O_3 \) plus other impurities. A 752.0 g sample of impure iron ore is heated with excess carbon, producing 453.0 g of pure iron by the following reaction:

\[
Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(s) + 3CO(g)
\]

What is the mass percent of \( Fe_2O_3 \) in the impure iron ore sample? Assume that \( Fe_2O_3 \) is the only source of iron and that the reaction is 100% efficient.

148. Commercial brass, an alloy of \( Zn \) and \( Cu \), reacts with hydro-chloric acid as follows:

\[
Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)
\]

(\( Cu \) does not react with HCl.) When 0.5065 g of a certain brass alloy is reacted with excess HCl, 0.0985 g \( ZnCl_2 \) is eventually isolated.

a. What is the composition of the brass by mass?

b. How could this result be checked without changing the above procedure?

149. Vitamin \( A \) has a molar mass of 286.4 g/mol and a general molecular formula of \( C_xH_yE \), where \( E \) is an unknown element. If vitamin \( A \) is 83.86% C and 10.56% H by mass, what is the molecular formula of vitamin \( A \)?

150. You have seven closed containers, each with equal masses of chlorine gas (\( Cl_2 \)). You add 10.0 g of sodium to the first sample, 20.0 g of sodium to the second sample, and so on (adding 70.0 g of sodium to the seventh sample).
Sodium and chlorine react to form sodium chloride according to the equation

\[ 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \]

After each reaction is complete, you collect and measure the amount of sodium chloride formed. A graph of your results is shown below.

Answer the following questions:

a. Explain the shape of the graph.

b. Calculate the mass of \( \text{NaCl} \) formed when 20.0 g of sodium is used.

c. Calculate the mass of \( \text{Cl}_2 \) in each container.

d. Calculate the mass of \( \text{NaCl} \) formed when 50.0 g of sodium is used.

e. Identify the leftover reactant, and determine its mass for parts b and d above.

151. A substance \( X_2Z \) has the composition (by mass) of 40.0% \( X \) and 60.0% \( Z \). What is the composition (by mass) of the compound \( XZ_2 \)?
Consider samples of phosphine (PH₃), water (H₂O), hydrogen sulfide (H₂S), and hydrogen fluoride (HF), each with a mass of 119g. Rank the compounds from the least to the greatest number of hydrogen atoms contained in the samples.

Calculate the number of moles for each compound in the following table.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Mass</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium phosphate</td>
<td>326.4g</td>
<td></td>
</tr>
<tr>
<td>Calcium nitrate</td>
<td>303.0g</td>
<td></td>
</tr>
<tr>
<td>Potassium chromate</td>
<td>141.6g</td>
<td></td>
</tr>
<tr>
<td>Dinitrogen pentoxide</td>
<td>406.3g</td>
<td></td>
</tr>
</tbody>
</table>

Arrange the following substances in order of increasing mass percent of nitrogen.

a. NO  
b. N₂O  
c. NH₃  
d. SNH

Para-cresol, a substance used as a disinfectant and in the manufacture of several herbicides, is a molecule that contains the elements carbon, hydrogen, and oxygen. Complete combustion of a 0.345g sample of p-cresol produced 0.983g carbon dioxide and 0.230g water. Determine the empirical formula for p-cresol.

A compound with molar mass 180.1g/mol has the following composition by mass:
### Table

<table>
<thead>
<tr>
<th>Element</th>
<th>Percentage</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>40.0%</td>
</tr>
<tr>
<td>H</td>
<td>6.70%</td>
</tr>
<tr>
<td>O</td>
<td>53.3%</td>
</tr>
</tbody>
</table>

Determine the empirical and molecular formulas of the compound.

157. Which of the following statements about chemical equations is(are) true?

a. When balancing a chemical equation, you can never change the coefficient in front of any chemical formula.

b. The coefficients in a balanced chemical equation refer to the number of grams of reactants and products.

c. In a chemical equation, the reactants are on the right and the products are on the left.

d. When balancing a chemical equation, you can never change the subscripts of any chemical formula.

e. In chemical reactions, matter is neither created nor destroyed so a chemical equation must have the same number of atoms on both sides of the equation.

158. Consider the following unbalanced chemical equation for the combustion of pentane ($C_5H_{12}$):

\[
C_5H_{12}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)
\]

If 20.4g of pentane are burned in excess oxygen, what mass of water can be produced, assuming 100% yield?

159. Sulfur dioxide gas reacts with sodium hydroxide to form sodium sulfite and water. The unbalanced chemical equation for this reaction is given below:

\[
SO_2(g) + NaOH(s) \rightarrow Na_2SO_3(s) + H_2O(l)
\]

Assuming you react 38.3g sulfur dioxide with 32.8g sodium hydroxide and assuming that the reaction goes to completion, calculate the mass of each product formed.
160. Gallium arsenide, $\text{GaAs}$, has gained widespread use in semiconductor devices that convert light and electrical signals in fiber-optic communications systems. Gallium consists of $60.\% ^{69}\text{Ga}$ and $40.\% ^{71}\text{Ga}$. Arsenic has only one naturally occurring isotope, $^{75}\text{As}$. Gallium arsenide is a polymeric material, but its mass spectrum shows fragments with the formulas $\text{GaAs}$ and $\text{Ga}_2\text{As}_2$. What would the distribution of peaks look like for these two fragments?

161. Consider the following data for three binary compounds of hydrogen and nitrogen:

<table>
<thead>
<tr>
<th></th>
<th>% H (by Mass)</th>
<th>% N (by Mass)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>17.75</td>
<td>82.25</td>
</tr>
<tr>
<td>II</td>
<td>12.58</td>
<td>87.42</td>
</tr>
<tr>
<td>III</td>
<td>2.34</td>
<td>97.66</td>
</tr>
</tbody>
</table>

When 1.00L of each gaseous compound is decomposed to its elements, the following volumes of $\text{H}_2(g)$ and $\text{N}_2(g)$ are obtained:

<table>
<thead>
<tr>
<th></th>
<th>$\text{H}_2(L)$</th>
<th>$\text{N}_2(L)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>1.50</td>
<td>0.50</td>
</tr>
<tr>
<td>II</td>
<td>2.00</td>
<td>1.00</td>
</tr>
<tr>
<td>III</td>
<td>0.50</td>
<td>1.50</td>
</tr>
</tbody>
</table>
162. Natural rubidium has the average mass of 85.4678u and is composed of isotopes $^{85}\text{Rb}$ (mass = 84.9117u) and $^{87}\text{Rb}$. The ratio of atoms $^{85}\text{Rb}/^{87}\text{Rb}$ in natural rubidium is 2.591. Calculate the mass of $^{87}\text{Rb}$.

163. A compound contains only carbon, hydrogen, nitrogen, and oxygen. Combustion of 0.157g of the compound produced 0.213g CO$_2$ and 0.0310gH$_2$O. In another experiment, it is found that 0.103g of the compound produces 0.0230gNH$_3$. What is the empirical formula of the compound? Hint: Combustion involves reacting with excess O$_2$. Assume that all the carbon ends up in CO$_2$ and all the hydrogen ends up in H$_2$O. Also assume that all the nitrogen ends up in the NH$_3$ in the second experiment.

164. Nitric acid is produced commercially by the Ostwald process, represented by the following equations:

$$4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$$
$$2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$$
$$3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)$$

What mass of NH$_3$ must be used to produce 1.0 x 10$^6$kgHNO$_3$ by the Ostwald process? Assume 100% yield in each reaction, and assume that the NO produced in the third step is not recycled.

165. When the supply of oxygen is limited, iron metal reacts with oxygen to produce a mixture of FeO and Fe$_2$O$_3$. In a certain experiment, 20.00g iron metal was reacted with 11.20g oxygen gas. After the experiment, the iron was totally consumed, and 3.24g oxygen gas remained. Calculate the amounts of FeO and Fe$_2$O$_3$ formed in this experiment.

166. A 9.780-g gaseous mixture contains ethane (C$_2$H$_6$) and propane (C$_3$H$_8$). Complete combustion to form carbon dioxide and water requires 1.120mole of oxygen gas. Calculate the mass percent of ethane in the original mixture.

167. Zinc and magnesium metal each reacts with hydrochloric acid to make chloride salts of the respective metals, and hydrogen gas. A 10.00-g mixture of zinc and magnesium produces 0.5171g of hydrogen gas upon being mixed with an excess of hydrochloric acid. Determine the percent magnesium by mass in the original mixture.

168. A gas contains a mixture of NH$_3(g)$ and N$_2$H$_4(g)$, both of which react with O$_2(g)$ to form NO$_2(g)$ and H$_2$O(g). The gaseous mixture (with an initial mass
of 61.00g) is reacted with 10.00 moles O₂, and after the reaction is complete, 4.062 moles of O₂ remains. Calculate the mass percent of N₂H₄(g) in the original gaseous mixture.

169. Consider a gaseous binary compound with a molar mass of 62.09g/mol. When 1.39g of this compound is completely burned in excess oxygen, 1.21g of water is formed. Determine the formula of the compound. Assume water is the only product that contains hydrogen.

170. A 2.25-g sample of scandium metal is reacted with excess hydrochloric acid to produce 0.1502g hydrogen gas. What is the formula of the scandium chloride produced in the reaction?

171. In the production of printed circuit boards for the electronics industry, a 0.60-mm layer of copper is laminated onto an insulating plastic board. Next, a circuit pattern made of a chemically resistant polymer is printed on the board. The unwanted copper is removed by chemical etching, and the protective polymer is finally removed by solvents. One etching reaction is

\[ \text{Cu(NH}_3\text{)}_4\text{Cl}_2(aq) + 4\text{NH}_3(aq) + \text{Cu}(s) \rightarrow 2\text{Cu(NH}_3\text{)}_4\text{Cl}(aq) \]

A plant needs to manufacture 10,000 printed circuit boards, each 8.0 × 16.0cm in area. An average of 80% of the copper is removed from each board (density of copper = 8.96g/cm³). What masses of Cu(NH₃)₄Cl₂ and NH₃ are needed to do this? Assume 100% yield.

172. The aspirin substitute, acetaminophen (C₈H₉O₂N), is produced by the following three-step synthesis:

I. C₆H₅O₃N(s) + 3H₂(g) + HCl(aq) → C₆H₈ONCl(s) + 2H₂O(l)

II. C₆H₈ONCl(s) + NaOH(aq) → C₆H₇ON(s) + H₂O(l) + NaCl(aq)

III. C₆H₇ON(s) + C₄H₆O₃(l) → C₅H₉O₂N(s) + HC₃H₃O₂(l)

The first two reactions have percent yields of 87% and 98% by mass, respectively. The overall reaction yields 3 moles of acetaminophen product for every 4 moles of C₆H₅O₃N reacted.

a. What is the percent yield by mass for the overall process?

b. What is the percent yield by mass of Step III?

173. An element X forms both a dichloride (XCl₂) and a tetrachloride (XCl₄).
Treatment of 10.00g XCl₂ with excess chlorine forms 12.55g XCl₄. Calculate the atomic mass of X, and identify X.

174. When M₂S₃(s) is heated in air, it is converted to MO₂(s). A 4.000-g sample of M₂S₃(s) shows a decrease in mass of 0.277g when it is heated in air. What is the average atomic mass of M?

175. When aluminum metal is heated with an element from Group 6A of the periodic table, an ionic compound forms. When the experiment is performed with an unknown Group 6A element, the product is 18.56%Al by mass. What is the formula of the compound?

176. Consider a mixture of potassium chloride and potassium nitrate that is 43.2% potassium by mass. What is the percent KCl by mass of the original mixture?

177. Ammonia reacts with O₂ to form either NO(g) or NO₂(g) according to these unbalanced equations:

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

In a certain experiment 2.00 moles of NH₃(g) and 10.00 moles of O₂(g) are contained in a closed flask. After the reaction is complete, 6.75 moles of O₂(g) remains. Calculate the number of moles of NO(g) in the product mixture. *(Hint: You cannot do this problem by adding the balanced equations because you cannot assume that the two reactions will occur with equal probability.)*

178. You take 1.00g of an aspirin tablet (a compound consisting solely of carbon, hydrogen, and oxygen), burn it in air, and collect 2.20g CO₂ and 0.400g H₂O. You know that the molar mass of aspirin is between 170 and 190g/mol.

Reacting 1 mole of salicylic acid with 1 mole of acetic anhydride (C₄H₆O₃) gives you 1 mole of aspirin and 1 mole of acetic acid (C₂H₄O₂). Use this information to determine the molecular formula of salicylic acid.

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Chapter 3: Stoichiometry Integrative Problems
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**For Review**

**Integrative Problems**

These problems require the integration of multiple concepts to find the solutions.
179. With the advent of techniques such as scanning tunneling microscopy, it is now possible to “write” with individual atoms by manipulating and arranging atoms on an atomic surface.

a. If an image is prepared by manipulating iron atoms and their total mass is $1.05 \times 10^{-20}$ g, what number of iron atoms were used?

b. If the image is prepared on a platinum surface that is exactly 20 platinum atoms high and 14 platinum atoms wide, what is the mass (grams) of the atomic surface?

c. If the atomic surface were changed to ruthenium atoms and the same surface mass as determined in part b is used, what number of ruthenium atoms is needed to construct the surface?

180. Tetrodotoxin is a toxic chemical found in fugu pufferfish, a popular but rare delicacy in Japan. This compound has an LD$_{50}$ (the amount of substance that is lethal to 50% of a population sample) of 10 $\mu$g per kg of body mass. Tetrodotoxin is 41.38% carbon by mass, 13.16% nitrogen by mass, and 5.37% hydrogen by mass, with the remaining amount consisting of oxygen. What is the empirical formula of tetrodotoxin? If three molecules of tetrodotoxin have a mass of $1.59 \times 10^{-21}$ g, what is the molecular formula of tetrodotoxin? What number of molecules of tetrodotoxin would be the LD$_{50}$ dosage for a person weighing 165 lb?

181. An ionic compound MX$_3$ is prepared according to the following unbalanced chemical equation.

$$M + X_2 \rightarrow MX_3$$

A 0.105-g sample of X$_2$ contains $8.92 \times 10^{20}$ molecules. The compound MX$_3$ consists of 54.47% X by mass. What are the identities of M and X, and what is the correct name for MX$_3$? Starting with 1.00 g each of M and X$_2$, what mass of MX$_3$ can be prepared?

182. The compound As$_2$I$_4$ is synthesized by reaction of arsenic metal with arsenic triiodide. If a solid cubic block of arsenic ($d = 5.72$ g/cm$^3$) that is 3.00 cm on edge is allowed to react with $1.01 \times 10^{24}$ molecules of arsenic triiodide, what mass of As$_2$I$_4$ can be prepared? If the percent yield of As$_2$I$_4$ was 75.6%, what mass of As$_2$I$_4$ was actually isolated?
For Review

Marathon Problems

These problems are designed to incorporate several concepts and techniques into one situation.

Marathon Problems can be used in class by groups of students to help facilitate problem-solving skills.

183. A 2.077-g sample of an element, which has an atomic mass between 40 and 55, reacts with oxygen to form 3.708g of an oxide. Determine the formula of the oxide (and identify the element).

184. Consider the following balanced chemical equation:

\[ A + 5B \rightarrow 3C + 4D \]

a. Equal masses of A and B are reacted. Complete each of the following with either “A is the limiting reactant because _______; “B is the limiting reactant because _______;” or “we cannot determine the limiting reactant because _______.”

i. If the molar mass of A is greater than the molar mass of B, then

ii. If the molar mass of B is greater than the molar mass of A, then

b. The products of the reaction are carbon dioxide (C) and water (D). Compound A has a similar molar mass to carbon dioxide. Compound B is a diatomic molecule. Identify compound B, and support your answer.

c. Compound A is a hydrocarbon that is 81.71\% carbon by mass. Determine its empirical and molecular formulas.

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1. When a sample of an unknown element is vaporized and injected into a mass spectrometer, the results shown below are obtained. Use these data to estimate the average atomic mass of this element.

![Mass Spectrometer Graph]

(A) 117amu
(B) between 117 and 118amu
(C) between 118 and 119amu
(D) between 119 and 120amu

2. An unnamed element with an atomic number of 130 is vaporized and injected into a mass spectrometer. The results are shown below. Use these data to calculate the average atomic mass of this element.

![Mass Spectrometer Graph]

(A) 320amu
(B) 321amu
(C) 322amu
(D) 324amu

3. Phosphorus can be produced by reacting calcium phosphate, silicon dioxide,
and carbon at high temperatures according to the following equation:

\[ \text{Ca}_3(\text{PO}_4)_2(s) + \text{SiO}_2(s) + \text{C}(s) \rightarrow \text{CaSiO}_3(s) + \text{P}_4(s) + \text{CO}(g) \]

What is the sum of the coefficients when the equation is balanced in standard form (lowest multiple whole numbers)?

(A) 16  
(B) 25  
(C) 32  
(D) 35

4. Nitroglycerine, \( \text{C}_3\text{H}_5\text{N}_3\text{O}_9 \), explodes with tremendous force due to the numerous gaseous products. The unbalanced equation for the explosion of nitroglycerine is shown below. What is the coefficient of the \( \text{CO}_2(g) \) when the equation is balanced using the lowest whole number coefficients?

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

(A) 8  
(B) 10  
(C) 12  
(D) 14

5. The simplest formula for a compound made from element \( X \) (molar mass = 79.0 g/mol) that is 21.0% nitrogen by mass is

(A) \( \text{XN} \)  
(B) \( \text{XN}_2 \)  
(C) \( \text{X}_2\text{N}_3 \)  
(D) \( \text{X}_3\text{N}_2 \)

6. When 78.0 g of aluminum hydroxide, \( \text{Al(OH)}_3 \) (molar mass = 78.0 g/mol), reacts with 49.0 g sulfuric acid, \( \text{H}_2\text{SO}_4 \) (molar mass = 98.1 g/mol), what mass of water is produced? The equation is

\[ 2\text{Al(OH)}_3(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 6\text{H}_2\text{O}(l) \]

(A) 18.0 g
7. What mass of aluminum is required if 40.0 g of iron(III) oxide, Fe₂O₃, is to be completely consumed as shown in the following reaction?

\[
\text{Fe₂O₃(s) + 2Al(s) → 2Fe(l) + Al₂O₃(s)}
\]

(A) 13.5 g
(B) 27.0 g
(C) 40.0 g
(D) 67.0 g

8. The mineral magnesite contains magnesium carbonate, \(\text{MgCO₃}\) (molar mass = 84 g/mol), and other impurities. When a 1.26 g sample of magnesite was dissolved in hydrochloric acid, 0.22 g of \(\text{CO₂}\) was generated. If the magnesite contained no carbonate other than \(\text{MgCO₃}\), what was the percent \(\text{MgCO₃}\) by mass in the magnesite? The equation is

\[
\text{MgCO₃(s) + 2HCl(aq) → MgCl₂(aq) + H₂O(l) + CO₂(g)}
\]

(A) 25%
(B) 33%
(C) 50.0%
(D) 67%

9. A compound is analyzed and found to contain 1.30 moles of \(\text{Na}\), 0.65 moles of \(\text{Te}\), and 2.60 moles of \(\text{O}\). What is the simplest formula of this compound?

(A) \(\text{NaTeO₂}\)
(B) \(\text{Na₂TeO₄}\)
(C) \(\text{Na₂Te₂O}\)
(D) \(\text{Na₂Te₄O}\)

10. The average atomic mass of a chlorine atom is 35.45 amu. Chlorine consists of two isotopes, \(^{35}\text{Cl}\) and \(^{37}\text{Cl}\). Which of the following most closely approximates the relative abundance of these isotopes?
11. What is the sum of the coefficient integers when the following equation is balanced in standard form (lowest multiple whole numbers)?

\[ \text{FeCr}_2\text{O}_7(s) + \text{K}_2\text{CO}_3(s) + \text{O}_2(g) \rightarrow \text{K}_2\text{CrO}_4(s) + \text{Fe}_2\text{O}_3(s) + \text{CO}_2(g) \]

(A) 9  
(B) 15  
(C) 24  
(D) 31

12. When iron(III) oxide is heated with carbon, the products are iron metal and carbon monoxide. Balance the equation in standard form and determine the sum of the coefficients.

(A) 4  
(B) 5  
(C) 6  
(D) 9

13. Methanol (CH\textsubscript{3}OH) can react with oxygen gas to produce formaldehyde (H\textsubscript{2}CO) and water. How much formaldehyde can be produced by reacting 4.0 moles of methanol with 4.0 moles of oxygen gas?

(A) 2.0 moles  
(B) 4.0 moles  
(C) 6.0 moles  
(D) 8.0 moles

14. Which of the following compounds contains the greatest percent by mass of nitrogen?

(A) NH\textsubscript{3}
15. Substance $AB_2$ is 60.0% $A$ by mass. What is the percent $A$ by mass for substance $AB$?

(A) 20.0%

(B) 40.0%

(C) 60.0%

(D) 75.0%

16. In a Bunsen burner, methane reacts with oxygen to produce carbon dioxide and water, as shown in the balanced equation below:

$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$$

What mass of water can be produced from 16 grams of methane and excess oxygen?

(A) 8.0 g

(B) 16 g

(C) 18 g

(D) 36 g

17. Which of the following reaction mixtures would produce the greatest amount of product, assuming all went to completion? Each involves the reaction symbolized by the equation:

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

(A) 2 mol $\text{N}_2$ and 4 mol $\text{H}_2$

(B) 1 mol $\text{N}_2$ and 5 mol $\text{H}_2$

(C) 4 mol $\text{N}_2$ and 2 mol $\text{H}_2$

(D) 5 mol $\text{N}_2$ and 1 mol $\text{H}_2$