Chapter 3
Stoichiometry

In This Chapter…

As you have learned in previous chapters, much of chemistry involves using macroscopic measurements to deduce what happens between atoms and molecules. We will now explore the chemical counting unit that links the atomic and macroscopic scales, the mole. The mole will allow us to study in greater detail chemical formulas and chemical reactions. Specifically, we will investigate stoichiometry, the relationship between quantities of materials in chemical reactions. In the next chapter we will expand our study of stoichiometry to include different types of chemical reactions and focus on reactions that take place in water.

Chapter Outline

Section 3.1 The Mole and Molar Mass

Section Outline

3.1a Avogadro's Number
3.1b Molar Mass
Section Summary Assignment

As you learned in Chapter 1, atoms are so small and have such small masses that any amount of atoms we would work with would be very hard to count. For example, a piece of aluminum about the size of pencil eraser contains approximately 2×10^{22} aluminum atoms! Chemists use a special unit when counting the numbers of atoms or molecules in a sample that has been measured at the macroscopic level. This unit is called the mole, and it is based on quantities that we work with easily in the laboratory. For example, the eraser-sized sample of aluminum that contains about 2×10^{22} Al atoms can also be described as containing 0.03 moles of Al atoms.

Opening Exploration 3.1 Avogadro's Number

These samples of copper, aluminum and sulfur all contain the same number of atoms.

3.1a Avogadro's Number

The **mole** (abbreviated **mol**) is the unit chemists use when counting numbers of atoms or molecules in a sample. The number of particles (atoms, molecules, or other objects) in one mole is equal to the number of atoms in exactly 12 g of carbon-12. This number of particles is called **Avogadro's number** (***N***) and has a value of 6.0221415×10^{23}. In most cases we will use 6.022×10^{23} or 6.02×10^{23} for Avogadro's number. One mole of any element contains 6.0221415×10^{23} atoms of that element, and one mole of a molecular compound contains 6.0221415×10^{23} molecules of that compound. Avogadro's number is an extremely large number, as it must be to connect tiny atoms to the macroscopic world. Using the mole counting unit to measure something on the macroscopic scale demonstrates just how big Avogadro's number is. For example, one mole of pencil erasers would cover the Earth's surface to a depth of about 500 meters. Interactive Figure 3.1.1 shows one mole quantities of some elements and compounds.

Interactive Figure 3.1.1 Recognize How the Mole Connects Macroscopic and Atomic Scales
One mole quantities of (from left to right) the elements copper, aluminum, sulfur, and oxygen and the compounds potassium dichromate, water, and copper(II) chloride dihydrate.

Avogadro's number can be used to convert between moles and the number of particles in a sample, as shown in the following example.

EXAMPLE PROBLEM: Convert Between Moles and Numbers of Atoms

A sample of titanium contains 8.98×10^{25} Ti atoms. What amount of Ti, in moles, does this represent?

SOLUTION:

You are asked to calculate the amount (moles) of Ti in a given sample of the metal.

You are given the number of atoms of Ti in the sample.

Use the equality 1 mol = 6.022×10^{23} atoms to create a conversion factor that converts from units of atoms to units of moles.

$$8.98 \times 10^{25} \text{ Ti atoms} \times \frac{1 \text{ mol Ti}}{6.022 \times 10^{23} \text{ Ti atoms}} = 149 \text{ mol Ti}$$

Is your answer reasonable? The sample contains more than Avogadro's number of Ti atoms and therefore contains more than 1 mol of Ti atoms.

OWL: 3.1.1T: Tutorial Assignment
3.1.1: Mastery Assignment

3.1b Molar Mass

Molar mass is the mass, in grams, of one mole of of a substance. The molar mass of an element is the mass in grams of one mole of atoms of the element. It is related to the atomic weight of an element, as shown here:

$$1\ {}^{12}\text{C atom} = 12\ \text{u}$$

$$6.022 \times 10^{23}\ (1\ \text{mole})\ {}^{12}\text{C atoms} = 12\ \text{grams}$$

Thus, one mole of an element has a mass in grams equal to its atomic weight in atomic mass units (u). For example, according to the periodic table, the element magnesium has an atomic weight of 24.31 u, and one mole of magnesium atoms has a mass of 24.31 g.

Just as Avogadro's number can be used to convert between moles and the number of particles in a sample, molar mass can be used to convert between moles and the mass in grams of a sample, as shown in the following example. Note that the number of significant figures in molar mass and Avogadro's number can vary. In calculations, include one more significant figure than the data given in the problem so that they do not limit the number of significant figures in the answer.

EXAMPLE PROBLEM: Convert Between Mass and Moles of an Element

What amount of oxygen, in moles, does 124 g O represent?

SOLUTION:

You are asked to calculate the amount (in moles) of oxygen atoms in a given sample.

You are given the mass of the oxygen sample.

Use the molar mass of oxygen (1 mol O = 16.00 g) to create a conversion factor that converts mass (in grams) to amount (in mol) of oxygen.

$$124\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g}} = 7.75\ \text{mol O}$$

Notice that the units of grams cancel, leaving the answer in units of mol. Also, note that the molar mass of oxygen has four significant figures, one more than the number of significant figures in the data given in the problem.

Is your answer reasonable? The mass of oxygen in the sample is greater than the molar mass of oxygen, so there is more than 1 mol O in the sample.

OWL: 3.1.2T: Tutorial Assignment
3.1.2: Mastery Assignment

We can now use Avogadro's number and molar mass to relate mass, moles, and atoms of an element, as shown in the following example.

EXAMPLE PROBLEM: Convert Between Mass, Moles, and Atoms of an Element

How many boron atoms are there in a 77.8-g sample of elemental boron?

SOLUTION:

You are asked to calculate the number of boron atoms in a given sample.

You are given the mass of the boron sample.

This calculation involves two conversion factors: the molar mass of boron to convert mass (in grams) to amount (in mol) and Avogadro's number to convert amount (in mol) to amount (atoms).

$$77.8 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ atoms B}}{1 \text{ mol B}} = 4.33 \times 10^{24} \text{ atoms B}$$

Notice that the units of grams and mol cancel, leaving the answer in units of atoms. Also, note that the molar mass and Avogadro's number have four significant figures, one more than the number of significant figures in the data given in the problem.

Is your answer reasonable? The mass of the boron sample is greater than the molar mass of boron, so the number of boron atoms in the sample is greater than Avogadro's number.

OWL: **3.1.3T: Tutorial Assignment**
3.1.3: Mastery Assignment

The molar mass of a compound is the mass in grams of one mole of the compound. It is numerically equal to the compound's formula (or molecular) weight. **Formula weight** is the sum of the atomic weights of the elements that make up a substance multiplied by the number of atoms of each element in the formula for the substance. For substances that exist as individial molecules, the formula weight is called the **molecular weight**. For example, one molecule of water, H_2O, is made up of two hydrogen atoms and one oxygen atom. Therefore, the mass of 1 mol of H_2O molecules (the molar mass of H_2O) is equal to the mass of 2 mol of hydrogen atoms plus the mass of 1 mol of oxygen atoms.

$$\text{Molar mass } H_2O = (2 \text{ mol H})\left(\frac{1.01 \text{ g}}{1 \text{ mol H}}\right) + (1 \text{ mol O})\left(\frac{16.00 \text{ g}}{1 \text{ mol O}}\right)$$

$$\text{Molar mass } H_2O = 18.02 \text{ g/mol}$$

In this text we will generally report the molar mass of an element or compound to two decimal places, unless more significant figures are required in a calculation.

EXAMPLE PROBLEM: Determine the Molar Mass of a Compound

Calculate the molar mass for each of the following compounds:

(a) 2-Propanol, $CH_3CH(OH)CH_3$

(b) Iron(II) phosphate, $Fe_3(PO_4)_2$

SOLUTION:

You are asked to calculate the molar mass of a compound.

You are given the compound's formula.

Add the molar masses for the constituent elements, taking into account the number of each element present in the compound.

(a) One mole of 2-propanol, $CH_3CH(OH)CH_3$, contains three moles of C, eight moles of H, and one mole of O.

C: $3 \text{ mol C} \times \dfrac{12.01 \text{ g}}{1 \text{ mol C}} = 36.03 \text{ g C}$

H: $8 \text{ mol H} \times \dfrac{1.01 \text{ g}}{1 \text{ mol H}} = 8.08 \text{ g H}$

O: $1 \text{ mol O} \times \dfrac{16.00 \text{ g}}{1 \text{ mol O}} = 16.00 \text{ g O}$

Molar mass of 2-propanol = 36.03 g C + 8.08 g H + 16.00 g O = 60.11 g/mol

(b) One mole of iron(II) phosphate, $Fe_3(PO_4)_2$, contains three moles of Fe, two moles of P, and eight moles of O.

Fe: $3 \text{ mol Fe} \times \dfrac{55.85 \text{ g}}{1 \text{ mol Fe}} = 167.6 \text{ g Fe}$

P: $2 \text{ mol P} \times \dfrac{30.97 \text{ g}}{1 \text{ mol P}} = 61.94 \text{ g P}$

O: $8 \text{ mol O} \times \dfrac{16.00 \text{ g}}{1 \text{ mol O}} = 128.0 \text{ g O}$

Molar mass iron(II) phosphate = 167.6 g Fe + 61.94 g P + 128.0 g O = 357.5 g/mol

3.1.4T: Tutorial Assignment

3.1.4: Mastery Assignment

The molar mass of a compound can be used to convert between moles and the mass in grams of a sample, as shown in the following example.

EXAMPLE PROBLEM: Convert Between Mass and Moles of a Compound

Use the molar mass of 2-propanol, 60.11 g/mol, to calculate the amount of 2-propanol present in a 10.0-g sample of the alcohol.

SOLUTION:

You are asked to calculate the amount (mol) of compound present in a given sample.

You are given the mass of the sample and the molar mass of the compound.

Use the molar mass of 2-propanol to create a conversion factor that converts mass (in grams) to amount (in mol) of 2-propanol.

$$10.0\text{ g }CH_3CH(OH)CH_3 \times \frac{1\text{ mol }CH_3CH(OH)CH_3}{60.11\text{ g}} = 0.166\text{ mol }CH_3CH(OH)CH_3$$

Note that the molar mass used in the calculation has four significant figures, one more than the data given in the problem.

Is your answer reasonable? The mass of the sample is less than the molar mass of the compound, so the amount of compound present is less than 1 mole.

OWL: **3.1.5T: Tutorial Assignment**
3.1.5: Mastery Assignment

3.2. Stoichiometry and Compound Formulas

Section Outline

Stoichiometry can be used to determine the chemical formula of a compound by studying the relative amounts of elements present or can be used to study the relative amounts of compounds that are consumed and produced during a chemical reaction. At the heart of stoichiometry is the mole, the relationship between macroscopic amounts and the atomic level.

Opening Exploration 3.2 Compounds and Moles

[UN Figure 3.2]

3.2a Element Composition

Stoichiometry is the study of the relationship between relative amounts of substances. The formula of a compound provides information about the relative amount of each element present in either one molecule of the compound or one mole of the compound. For example, one molecule of acetic acid, CH_3CO_2H, contains two atoms of oxygen and one mole of acetic acid contains 2 mol of oxygen atoms. When working with ionic and other types of nonmolecular compounds, the compound formula is still used to describe the stoichiometry of a compound. For example, the ionic compound calcium chloride, $CaCl_2$, is not made up of $CaCl_2$ molecules but rather one mole of $CaCl_2$ contains 1 mol Ca^{2+} ions and 2 mol Cl^- ions.

The compound formula can be used to determine the amount of an element present in a sample of compound, as shown in the following example.

EXAMPLE PROBLEM: Use Compound Formulas to Determine Element Composition

A sample of acetic acid, CH_3CO_2H, contains 2.50 mol of the compound. Determine the amount (in mol) of each element present and the number of atoms of each element present in the sample.

SOLUTION:

You are asked to calculate the amount (mol) and number of atoms of each element present in a given sample of a compound.

You are given the amount of sample (mol) and the chemical formula of the compound.

Use the compound formula to create conversion factors that relate the amount (in mol) of each element to one mole of the compound

$$2.50 \text{ mol } CH_3CO_2H \times \frac{2 \text{ mol C}}{1 \text{ mol } CH_3CO_2H} = 5.00 \text{ mol C}$$

$$2.50 \text{ mol } CH_3CO_2H \times \frac{4 \text{ mol H}}{1 \text{ mol } CH_3CO_2H} = 10.0 \text{ mol H}$$

$$2.50 \text{ mol } CH_3CO_2H \times \frac{2 \text{ mol O}}{1 \text{ mol } CH_3CO_2H} = 5.00 \text{ mol O}$$

Use Avogadro's number and the amount of each element present to determine the number of atoms of each element present in the 2.50-mol sample of acetic acid.

$$5.00 \text{ mol C} \times \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C}} = 3.01 \times 10^{24} \text{ C atoms}$$

$$10.0 \text{ mol H} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} = 6.02 \times 10^{24} \text{ C atoms}$$

$$5.00 \text{ mol O} \times \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O}} = 3.01 \times 10^{24} \text{ O atoms}$$

Is your answer reasonable? The sample contains more than one mole of the compound, so it also contains more than one mole of each element and more than Avogadro's number of atoms of each element in the compound

OWL: **3.2.1T: Tutorial Assignment**
3.2.1: Mastery Assignment

3.2b Percent Composition

The mole-to-mole relationships in a chemical formula can also be used to determine the **percent composition** of an element in a compound. The percent composition of an element in a compound is the mass of an element present in exactly 100 g of a compound and is calculated using the following equation:

$$\%\ \text{element} = \frac{\text{(number of atoms of element in formula)(molar mass of element)}}{\text{mass of 1 mol of compound}} \times 100\% \quad (3.1)$$

For example, the percent composition of hydrogen in water, H_2O, is calculated as follows:

$$\%\ \text{H in H}_2\text{O} = \frac{2\ \text{mol H}\left(\dfrac{1.01\ \text{g}}{1\ \text{mol H}}\right)}{1\ \text{mol H}_2\text{O}\left(\dfrac{18.02\ \text{g}}{1\ \text{mol H}_2\text{O}}\right)} \times 100\% = 11.2\%\ \text{H}$$

Notice that percent composition is calculated by using the mole-to-mole ratio in the chemical formula of water (2 mol H in 1 mol H_2O) and converting it to a mass ratio using molar mass. Because water is made up of only two elements and the sum of all the percent composition values must equal 100%, the percent composition of oxygen in water is

$$\%\ \text{O in H}_2\text{O} = 100\% - \%\ \text{H} = 100\% - 11.2\% = 88.8\%\ \text{O}$$

EXAMPLE PROBLEM: Calculate Percent Composition from a Compound Formula

Determine the percent composition of each element in potassium permanganate, $KMnO_4$, a compound used as an antiseptic in some countries.

SOLUTION:

You are asked to calculate the percent composition of each element in a compound.

You are given the compound formula.

First calculate the molar mass of $KMnO_4$, then use it along with the chemical formula to calculate the percent composition of each element.

$$\text{Molar mass KMnO}_4 = (1\ \text{mol K})\left(\frac{39.10\ \text{g}}{1\ \text{mol K}}\right) + (1\ \text{mol Mn})\left(\frac{54.94\ \text{g}}{1\ \text{mol Mn}}\right) + (4\ \text{mol O})\left(\frac{16.00\ \text{g}}{1\ \text{mol O}}\right)$$

Molar mass $KMnO_4$ = 158.0 g/mol

$$\%\ \text{K in KMnO}_4 = \frac{1\ \text{mol K}\left(\dfrac{39.10\ \text{g}}{1\ \text{mol K}}\right)}{158.0\ \text{g KMnO}_4} \times 100\% = 24.75\%\ \text{K}$$

$$\%\ \text{Mn in KMnO}_4 = \frac{1\ \text{mol Mn}\left(\dfrac{54.94\ \text{g}}{1\ \text{mol Mn}}\right)}{158.0\ \text{g KMnO}_4} \times 100\% = 34.77\%\ \text{Mn}$$

$$\%\ \text{O in KMnO}_4 = \frac{4\ \text{mol O}\left(\dfrac{16.00\ \text{g}}{1\ \text{mol O}}\right)}{158.0\ \text{g KMnO}_4} \times 100\% = 40.51\%\ \text{O}$$

Is your answer reasonable? Each percentage should be less than 100%, and the sum of the percent composition values for the constituent elements must be equal to 100%. Notice that the oxygen percent composition could have been calculated from the % K and % Mn values.

% O in $KMnO_4$ = 100% – % K – % Mn = 100% – 24.75% – 34.77% = 40.48% O

The slight difference between the values is due to rounding during the calculation steps.

OWL: **3.2.2T: Tutorial Assignment**
3.2.2: Mastery Assignment

3.2c Empirical Formulas from Percent Composition

A common practice in the chemical laboratory is to determine the percent composition of each element in a compound and use that information to determine the formula that shows the simplest whole-number ratio of elements present in the compound, the **empirical formula** [**Flashback anchor**] of the compound. In a sense, this is the reverse of the process that is used to determine percent composition.

[**Flashback**] In Section 2.3 we discussed representations of covalent compounds, including empirical and molecular formulas.**[End Flashback.]**

[Flashback popup text] A molecular formula consists of symbols for each element present and a subscript number to identify the number of atoms of each element in the molecule. For example, the molecular formula of water is H_2O, indicating that a single water molecule consists of two hydrogen atoms and one oxygen atom. An empirical formula is the simplest whole-number ratio of elements in a compound. The empirical formula of water is the same as its molecular formula, H_2O. Hydrogen peroxide, however, has the molecular formula H_2O_2 and an empirical formula of HO. [/Flashback popup text]

The first step in the process is to understand that the percent composition of any element is equal to the mass of that element in exactly 100 g of compound. For example, a compound containing only phosphorus and chlorine is analyzed and found to contain 22.55% P and 77.45% Cl by mass. The percent composition of phosphorus can be written as

$$22.55\%\ \text{P} = \frac{22.55\ \text{g P}}{100\ \text{g compound}}$$

Thus, if we assume that we are working with a sample that has a mass of exactly 100 g, the mass of each element present is equal to the percent composition. For the compound containing only phosphorus and oxygen,

$$22.55\%\ \text{P} = 22.55\ \text{g P in 100 g compound}$$
$$77.45\%\ \text{Cl} = 77.45\ \text{g Cl in 100 g compound}$$

To determine the empirical formula of the compound, the mass of each element is converted to an amount in moles.

$$22.55\ \text{g P} \times \frac{1\ \text{mol P}}{30.974\ \text{g}} = 0.7280\ \text{mol P}$$

$$77.45\ \text{g Cl} \times \frac{1\ \text{mol Cl}}{35.453\ \text{g}} = 2.185\ \text{mol Cl}$$

Thus, this compound contains phosphorus and chlorine in a mole ratio of $P_{0.7280}Cl_{2.185}$. The empirical formula, however, shows the simplest whole-number ratio of elements. Dividing each amount by the smallest amount present produces integer subscripts without changing the relative amounts of each element present in the compound.

$$\frac{2.185 \text{ mol Cl}}{0.7280} = 3.001 \text{ mol Cl} \approx 3 \text{ mol Cl}$$

$$\frac{0.7280 \text{ mol P}}{0.7280} = 1.000 \text{ mol P} = 1 \text{ mol P}$$

Notice that amount of Cl can be rounded to the nearest whole number. The empirical formula of the compound, therefore, contains 3 mol of Cl and 1 mol of P, or PCl_3.

EXAMPLE PROBLEM: Use Percent Composition to Determine an Empirical Formula

A solid compound is found to contain K, S, and O with the following percent composition:

K: 41.09%
S: 33.70%
O: 25.22%

What is the empirical formula of this compound?

SOLUTION:

You are asked to determine the empirical formula for a compound.

You are given the percent composition of each element in the compound.

First, assume that you are given a 100-g sample, which would mean that the mass of each element in grams is equal to the percent composition value. Calculate the amount (in mol) of each element present, and then use the relative amounts to determine the empirical formula of the compound.

$$41.09 \text{ g K} \times \frac{1 \text{ mol K}}{39.098 \text{ g}} = 1.051 \text{ mol K}$$

$$33.70 \text{ g S} \times \frac{1 \text{ mol S}}{32.065 \text{ g}} = 1.051 \text{ mol S}$$

$$25.22 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g}} = 1.576 \text{ mol O}$$

To determine the simplest whole-number ratio of elements, divide each by the smallest value.

$$\frac{1.051 \text{ mol K}}{1.051} = 1.000 \text{ mol K} = 1 \text{ mol K}$$

$$\frac{1.051 \text{ mol S}}{1.051} = 1.000 \text{ mol S} = 1 \text{ mol S}$$

$$\frac{1.576 \text{ mol O}}{1.051} = 1.500 \text{ mol O} = 1.5 \text{ mol O}$$

Recall that the value 1.5 is the same as the fraction $^3/_2$. Multiply all three amounts by 2, the fraction denominator, to express the ratio of elements as whole numbers.

$$1 \text{ mol K} \times 2 = 2 \text{ mol K}$$

$$1 \text{ mol S} \times 2 = 2 \text{ mol S}$$

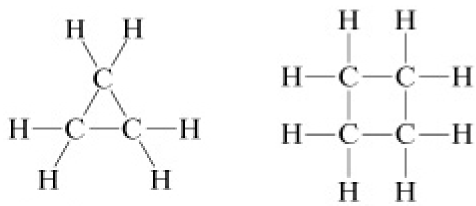
$$1.5 \text{ mol O} \times 2 = 3 \text{ mol O}$$

The empirical formula is therefore $K_2S_2O_3$.

OWL: **3.2.3T: Tutorial Assignment**
3.2.3: Mastery Assignment

3.2d Determining Molecular Formulas

The empirical formula of a compound gives the simplest whole-number ratio of atoms of each element present in a compound, but it may not be representative of the molecular formula of a compound. For example, the hydrocarbons cyclopropane and cyclobutane have the same empirical formula (CH_2) but different molecular formulas (C_3H_6 and C_4H_8, respectively). Ionic compounds and other compounds that do not exist as individual molecules are always represented using empirical formulas. For molecular compounds, additional information is needed to determine the molecular formula from an empirical formula.



Cyclopropane Cyclobutane

The **molecular formula** of a compound is a whole-number multiple of the empirical compound formula. For cyclobutane, the empirical formula (CH_2) has an molar mass of 14.03 g/mol and the molecular formula has a molar mass of 56.11 g/mol.

$$\frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \text{whole number}$$

$$\frac{56.11 \text{ g/mol}}{14.03 \text{ g/mol}} = 3.999 \approx 4$$

The whole-number multiple is equal to the number of empirical formula units in one molecule of the compound. For cyclobutane, there are four empirical formula units per molecule, and the molecular formula is $(CH_2)_4$ or C_4H_8. The molar mass of a compound must be determined, usually through additional experiments, in order to determine the molecular formula of a compound.

EXAMPLE PROBLEM: Use Percent Composition and Molar Mass to Determine Molecular Formula

Resorcinol is a compound used in the manufacture of fluorescent and leather dyes as well as to treat acne and other greasy skin conditions. Analysis of the compound showed that it is 65.45% C and 5.493% H, with oxygen accounting for the remainder. In a separate experiment, the molar mass of the compound was found to be 110.11 g/mol. Determine the molecular formula of resorcinol.

SOLUTION:

You are asked to determine the molecular formula of a compound.

You are given the percent composition of all but one element in the compound and the compound's molar mass.

First, determine the percent composition of oxygen.

$$\% \text{ O} = 100.00\% - \% \text{ C} - \% \text{ H} = 100.00\% - 65.45\% \text{ C} - 5.493\% \text{ H} = 29.06\% \text{ O}$$

Next, assume a 100-g sample and calculate the amount (in mol) of each element present.

$$65.45 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g}} = 5.449 \text{ mol C}$$

$$5.493 \text{ g H} \times \frac{1 \text{ mol H}}{1.0079 \text{ g}} = 5.450 \text{ mol H}$$

$$29.06 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g}} = 1.816 \text{ mol O}$$

To determine the simplest whole-number ratio of elements, divide each by the smallest value.

$$\frac{5.449 \text{ mol C}}{1.816} = 3.001 \text{ mol C} \approx 3 \text{ mol C}$$

$$\frac{5.450 \text{ mol H}}{1.816} = 3.001 \text{ mol H} \approx 3 \text{ mol H}$$

$$\frac{1.816 \text{ mol O}}{1.816} = 1.000 \text{ mol O} = 1 \text{ mol O}$$

The empirical formula is C_3H_3O. Use the molar mass of the empirical formula and the compound to determine the whole-number multiple that relates the empirical and molecular formulas.

$$\frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{110.11 \text{ g/mol}}{55.06 \text{ g/mol}} = 2$$

There are two empirical formula units in the molecular formula. The molecular formula of resorcinol is $(C_3H_3O)_2$ or $C_6H_6O_2$.

Is your answer reasonable? The molar mass of the compound should be a whole-number multiple of the empirical formula's molar mass. In this case, it is twice the empirical formula's molar mass, which suggests that the empirical formula was determined correctly.

OWL: **3.2.4T: Tutorial Assignment**

3.2.4: Mastery Assignment

3.2e Hydrated Compounds

A **hydrated ionic compound** is an ionic compound that has a well-defined amount of water trapped within the crystalline solid. The water associated with the compound is called the **water of hydration**. A hydrated compound formula includes the term $\cdot nH_2O$, where *n* is the number of moles of water incorporated into the solid per mole of ionic compound. Prefixes are used in naming hydrated compounds to indicate the number of waters of hydration.

Many solids used in the laboratory are hydrated. For example, reagent-grade copper(II) sulfate is usually provided as the hydrated compound $CuSO_4{\cdot}5H_2O$, copper(II) sulfate pentahydrate. Some common hydrated compounds and their uses are shown in Table 3.2.1. Notice that the molar mass of a hydrated compound includes the mass of the water of hydration.

Table 3.2 1 Some Common Hydrated Ionic Compounds

Molecular formula	Name	Molar mass (g/mol)	Common Name	Uses
$Na_2CO_3{\cdot}10H_2O$	Sodium carbonate decahydrate	286.14	Washing soda	Water softener
$Na_2S_2O_3{\cdot}5H_2O$	Sodium thiosulfate pentahydrate	248.18	Hypo	Photography
$MgSO_4{\cdot}7H_2O$	Magnesium sulfate heptahydrate	246.47	Epsom salt	Dyeing and tanning
$CaSO_4{\cdot}2H_2O$	Calcium sulfate dihydrate	172.17	Gypsum	Wallboard
$CaSO_4{\cdot}\frac{1}{2}H_2O$	Calcium sulfate hemihydrate	145.15	Plaster of Paris	Casts, molds
$CuSO_4{\cdot}5H_2O$	Copper(II) sulfate pentahydrate	249.68	Blue vitriol	Algicide, root killer

Heating a hydrated compound releases the water in the crystalline solid. For example, heating the compound $CuCl_2{\cdot}2H_2O$ releases two moles of water (in the form of water vapor) per mole of hydrated compound. As shown in the following example problem, we can determine the formula of a hydrated compound by performing this experiment in the laboratory.

EXAMPLE PROBLEM: Determine the formula of a hydrated compound.

A 32.86 g sample of a hydrate of $CoCl_2$ was heated thoroughly in a porcelain crucible until its weight remained constant. After heating, 17.93 g of the dehydrated compound remained in the crucible. What is the formula of the hydrate?

SOLUTION:

You are asked to determine the formula of an ionic hydrated compound.

You are given the mass of the hydrated compound and the mass of the compound when it has been dehydrated.

First, determine the mass of water lost when the hydrated compound was heated.

$32.86\text{ g} - 17.93\text{ g} = 14.93\text{ g } H_2O$

Next, calculate moles of water and moles of the dehydrated compound.

$$14.93\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.015\text{ g}} = 0.8288\text{ mol } H_2O$$

$$17.93\text{ g } CoCl_2 \times \frac{1\text{ mol } CoCl_2}{129.84\text{ g}} = 0.1381\text{ mol } CoCl_2$$

Finally, determine the simplest whole-number ratio of water to dehydrated compound.

$$\frac{0.8288\text{ mol } H_2O}{0.1381\text{ mol } CoCl_2} = \frac{6.001\text{ mol } H_2O}{1\text{ mol } CoCl_2} \approx \frac{6\text{ mol } H_2O}{1\text{ mol } CoCl_2}$$

The chemical formula for the hydrated compound is $CoCl_2{\cdot}6H_2O$.

OWL: 3.2.5T: Tutorial Assignment
3.2.5: Mastery Assignment

3.3 Stoichiometry and Chemical Reactions

Section Outline

Stoichiometry is used not only to determine the chemical formula of a compound but also to study the relative amounts of compounds involved in chemical reactions. Once again, we will use the mole to relate macroscopic amounts of material participating in a chemical reaction to the atomic level where atoms, ions, and molecules interact.

Opening Exploration 3.3 Making Alum from Aluminum

Aluminum like that used in common soda cans can be reacted to form the compound alum, $KAl(SO_4)_2{\cdot}12H_2O$.

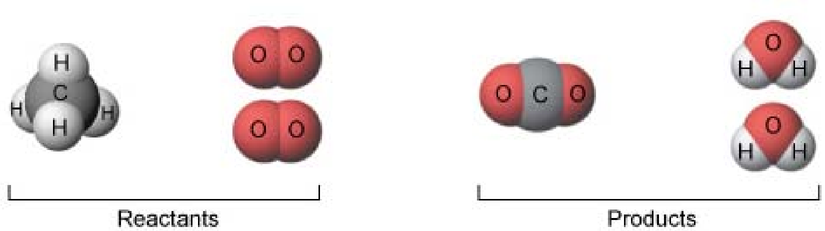
3.3a Chemical Reactions and Chemical Equations

Chemical reactions are represented by **chemical equations**, in which the reacting species (**reactants**) are shown to the left of a reaction arrow and the species produced during the reaction (**products**) appear to the right of the reaction arrow. In addition, the physical state of each reactant and product is often indicated using the symbols (g), (ℓ), (s), and (aq) for gas, liquid, solid, and aqueous (dissolved in water) solution, respectively. For example, the reaction of gaseous methane (CH_4) with gaseous molecular oxygen to produce gaseous carbon dioxide and water vapor is written as

$$\underbrace{CH_4(g) + 2\ O_2(g)}_{\text{reactants}} \rightarrow \underbrace{CO_2(g) + 2\ H_2O(g)}_{\text{products}}$$

Reading the equation, we say that one CH_4 molecule and two O_2 molecules react to form one CO_2 molecule and two H_2O molecules. The number that appears to the left of a compound formula in a balanced equation is called a **stoichiometric coefficient**; it indicates the relative number of molecules of that reactant or product in the reaction. (A stoichiometric coefficient of 1 is not written in a chemical equation.) Recall from Section 2.3 that the subscript to the right of each element symbol indicates the relative number of atoms of that element in the compound. Therefore, multiplying the stoichiometric coefficient of a compound by the subscript for each element in the compound formula gives the total number of atoms of each element that the compound contributes to the reactants or products.

The word *equation,* which means “equal on both sides,” reflects the fact that when a chemical reaction takes place, matter is neither created nor destroyed, in accordance with the **Law of Conservation of Matter**. This means that the products of a chemical reaction are made up of the same type and number of atoms found in the reactants—but rearranged into new compounds. In the reaction of methane with oxygen, for example, both the reactants and the products contain one carbon atom, four hydrogen atoms, and four oxygen atoms.

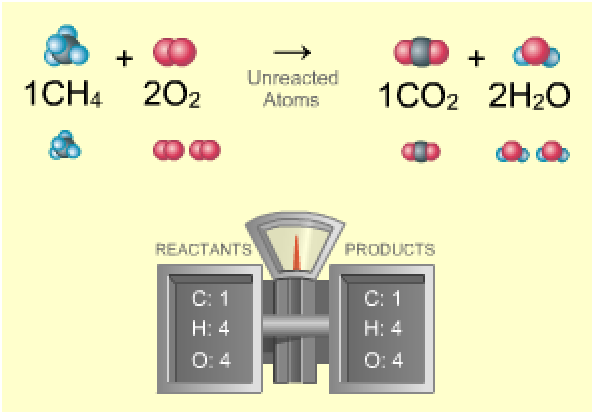


The reactants in this reaction are a CH_4 molecule and two O_2 molecules. When the reaction occurs, the chemical bonds between elements in the reactants are broken and new bonds are formed to make products. The same atoms are present both before and after the reaction, but the linkages between them have changed. When an equation meets the conditions of the Law of Conservation of Matter by having the same number of atoms of each element on both sides of the reaction arrow, the equation is **balanced**.

3.3b Balancing Chemical Equations

Very often we know what elements and compounds are involved in a reaction but not the relative amounts. To correctly describe the reaction, we must determine the relative amounts of the reactants and products involved in the reaction by balancing the equation (Interactive Figure 3.3.1).

Interactive Figure 3.3.1 Relate Conservation of Mass to Balanced Equations



The mass of reactants equals the mass of the products.

For example, consider the reaction that occurs in a gas grill when propane (C_3H_8) reacts with oxygen to form carbon dioxide and water. The unbalanced equation and the number of each element present in the reactants and product is

$$C_3H_8 + O_2 \rightarrow CO_2 + H_2O$$

3 C	1 C
8 H	2 H
2 O	3 O

The equation is not balanced because the number of atoms of each element present in the reactants is not equal to the number present in the products. The equation can be balanced by changing the coefficients to the left of each chemical species from 1 to a whole number that results in equal numbers of atoms of each element on both sides of the reaction arrow. It is important to note that the subscripts in a chemical formula are never changed when balancing an equation. A change in the subscript in a formula changes the chemical identify of the compound, not the amount of compound present in the reaction.

Although balancing chemical equations is usually done by trial and error, it is helpful to follow some general guidelines.

1. *If an element appears in more than one compound in the reactants or products, it is usually best to balance that element last.* In this case, we will begin by balancing carbon and hydrogen, balancing oxygen last because it appears in two compounds in the products.

2. As stated earlier, *only coefficients are changed when balancing equations, never the subscripts within a chemical formula.* For example, we cannot balance carbon by changing the formula of CO_2 to C_3O_2. Although this balances the number of carbon atoms, this change alters the chemical identity of the compound and makes the chemical equation invalid.

3. *Balanced chemical equations are written so that the coefficients are the lowest possible whole numbers.*

Balance carbon. In the unbalanced equation, there are three carbon atoms in the reactants and only one carbon atom in the products. Changing the coefficient in front of CO_2 from 1 to 3 balances the carbon atoms in the equation and also increases the number of oxygen atoms in the products from 3 to 7.

$$C_3H_8 + O_2 \rightarrow 3\ CO_2 + H_2O$$

3 C	3 C
8 H	2 H
2 O	7 O

Balance hydrogen. There are eight hydrogen atoms in the reactants and only two hydrogen atoms in the products. Each increase in the H_2O coefficient increases the number of hydrogen atoms by a factor of 2. Changing the coefficient in front of H_2O from 1 to 4 balances the number of hydrogen atoms in the equation and also increases the number of oxygen atoms in the products from 7 to 10.

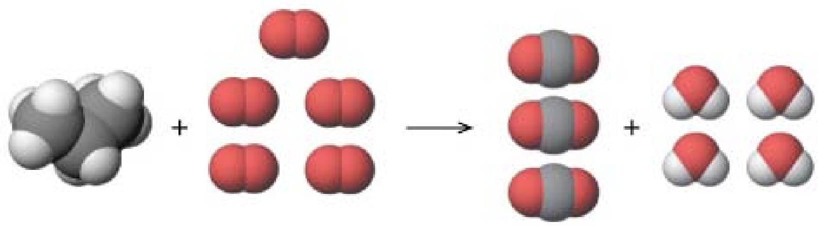
$$C_3H_8 + O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

3 C	3 C
8 H	8 H
2 O	10 O

Balance oxygen. There are 10 oxygen atoms in the products (6 in the three CO_2 molecules and 4 in the four H_2O molecules) and only 2 in the reactants. Each increase in the O_2 coefficient increases the number of oxygen atoms by a factor of 2, so changing the coefficient in front of O_2 from 1 to 5 balances the oxygen atoms in the equation.

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

3 C	3 C
8 H	8 H
10 O	10 O



The number of atoms of each element in the reactants and products is the same, so the equation is balanced. Finally, note that this equation can be balanced using a multiple of the coefficients above.

$$2\ C_3H_8 + 10\ O_2 \rightarrow 6\ CO_2 + 8\ H_2O$$

6 C	6 C
16 H	16 H
20 O	20 O

However, this is incorrect because it is possible to balance the equation using simpler coefficients.

EXAMPLE PROBLEM: Balance Equations

Balance the chemical equation for the neutralization of phosphoric acid by calcium hydroxide to form calcium phosphate and water. The unbalanced equation is $H_3PO_4 + Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$.

SOLUTION:

You are asked to balance a chemical equation.

You are given an unbalanced equation.

Notice that both hydrogen and oxygen appear in more than one compound in the reactants. Begin by balancing the other elements, phosphorus and calcium, before balancing hydrogen and oxygen.

$$H_3PO_4 + Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$$

5 H	2 H
1 P	2 P
6 O	9 O
1 Ca	3 Ca

Balance phosphorus. Change the coefficient in front of H_3PO_4 from 1 to 2 and note the change in H and O.

$$2\ H_3PO_4 + Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$$

8 H	2 H
2 P	2 P
10 O	9 O
1 Ca	3 Ca

Balance calcium. Change the coefficient in front of $Ca(OH)_2$ from 1 to 3 and note the change in H and O.

$$2\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$$

12 H	2 H
2 P	2 P
14 O	9 O
3 Ca	3 Ca

Balance hydrogen and oxygen. Changing the coefficient in front of H_2O from 1 to 6 (each increase in the H_2O coefficient adds two H atoms) balances both hydrogen and oxygen.

$$2\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6\ H_2O$$

12 H	12 H
2 P	2 P
14 O	14 O
3 Ca	3 Ca

The equation is balanced.

Is your answer reasonable? When balancing equations, it is a good idea to check the atom balance for all elements in the equation when you are finished. In this case, there are equal numbers of elements in both reactants and products, so the equation is balanced.

OWL: **3.3.1T: Tutorial Assignment**

3.3.1: Mastery Assignment

3.3c Reaction Stoichiometry

A balanced chemical equation shows the relative amounts of reactants and products involved in a chemical reaction on both the molecular and macroscopic scale. For example, the balanced equation for the reaction of methane with oxygen to form carbon dioxide and water describes the reaction at the atomic level and at the macroscopic level.

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$$

$$1\ CH_4 \text{ molecule} + 2\ O_2 \text{ molecules} \rightarrow 1\ CO_2 \text{ molecule} + 2\ H_2O \text{ molecules}$$

$$1 \text{ mol } CH_4 + 2 \text{ mol } O_2 \rightarrow 1 \text{ mol } CO_2 + 2 \text{ mol } H_2O$$

$$16.04 \text{ g } CH_4 + 64.00 \text{ g } O_2 \rightarrow 44.01 \text{ g } CO_2 + 36.03 \text{ g } H_2O$$

On the molecular level, one molecule of CH_4 reacts with two molecules of O_2 to produce one molecule of CO_2 and two molecules of H_2O. Scaling up to macroscopic amounts, the equation also represents the reaction of one mole of CH_4 (16.04 g) with two moles of O_2 (64.00 g) to form one mole of CO_2 (44.01 g) and two moles of H_2O (36.03 g). Reaction stoichiometry is the study of the relationships between the amount of reactants and products on the macroscopic scale, and balanced chemical equations are the key to understanding these relationships.

Consider the reaction between butane and oxygen to form carbon dioxide and water. The balanced equation is

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

According to the balanced equation, 8 mol of CO_2 are produced for every 2 mol of C_4H_{10} that react with oxygen. A typical reaction, however, might involve more or fewer than 2 mol of butane. The balanced equation gives us the information needed to determine the amount of CO_2 produced from any amount of C_4H_{10}.

The mole-to-mole relationships in a balanced chemical equation are used in the form of conversion factors to relate amounts of materials reacting or forming during a chemical reaction. For example, consider a butane lighter that contains 0.24 mol of C_4H_{10}. We can use the coefficients in the balanced equation to determine the amount of carbon dioxide produced in the reaction of this amount of butane with oxygen.

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

$$0.24 \text{ mol } C_4H_{10} \times \frac{8 \text{ mol } CO_2}{2 \text{ mol } C_4H_{10}} = 0.96 \text{ mol } CO_2$$

The ratio of coefficients from the balanced equation, called the **stoichiometric factor** (or *stoichiometric ratio*), relates the amount of one species to another. In this case, the stoichiometric factor relates the amount of carbon dioxide formed from the reaction of butane with oxygen. It is important to note that *stoichiometric factors are always created from coefficients in a balanced chemical equation.*

The amount of water produced in the reaction can also be determined. Again, a stoichiometric factor is created using the coefficients in the balanced equation that relate moles of H_2O and C_4H_{10}.

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

$$0.24 \text{ mol } C_4H_{10} \times \frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}} = 1.2 \text{ mol } H_2O$$

Stoichiometric relationships exist between all species in a balanced equation, not just between reactants and products. For example, if the reaction of a sample of butane with oxygen produces 2.8 mol of CO_2, it is possible to determine the amount of H_2O produced using a stoichiometric factor.

$$2.8 \text{ mol } CO_2 \times \frac{10 \text{ mol } H_2O}{8 \text{ mol } CO_2} = 3.5 \text{ mol } H_2O$$

EXAMPLE PROBLEM: Use Balanced Chemical Equations to Relate Amounts of Reactants and Products

The unbalanced equation for the reaction between magnesium nitride and sulfuric acid is shown here.

$$Mg_3N_2(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + (NH_4)_2SO_4(aq)$$

Balance the equation and determine the amount of H_2SO_4 consumed and the amounts of $MgSO_4$ and $(NH_4)_2SO_4$ produced when 3.5 mol of Mg_3N_2 reacts.

SOLUTION:

You are asked to balance a chemical equation and determine the amount of reactant consumed and products formed by a given amount of reactant.

You are given an unbalanced equation and the amount of one reactant consumed in the reaction.

Step 1. Write a balanced chemical equation.

$$Mg_3N_2(s) + 4\ H_2SO_4(aq) \rightarrow 3\ MgSO_4(aq) + (NH_4)_2SO_4(aq)$$

Step 2. Use the coefficients in the balanced equation to create a stoichiometric factor that will convert moles of Mg_3N_2 to moles of H_2SO_4 consumed.

$$3.5 \text{ mol } Mg_3N_2 \times \frac{4 \text{ mol } H_2SO_4}{1 \text{ mol } Mg_3N_2} = 14 \text{ mol } H_2SO_4 \text{ consumed}$$

Step 3. Use the coefficients in the balanced equation to create a stoichiometric factor that will convert moles of Mg_3N_2 to moles of $MgSO_4$ produced.

$$3.5 \text{ mol } Mg_3N_2 \times \frac{3 \text{ mol } MgSO_4}{1 \text{ mol } Mg_3N_2} = 10.5 \text{ mol } MgSO_4$$

Step 4. Use the coefficients in the balanced equation to create a stoichiometric factor that will convert moles of Mg_3N_2 to moles of $(NH_4)_2SO_4$ produced.

$$3.5 \text{ mol } Mg_3N_2 \times \frac{1 \text{ mol } (NH_4)_2SO_4}{1 \text{ mol } Mg_3N_2} = 3.5 \text{ mol } (NH_4)_2SO_4$$

Is your answer reasonable? According to the balanced equation, one mole of Mg_3N_2 reacts with four moles of H_2SO_4, producing three moles of $MgSO_4$ and one mole of $(NH_4)_2SO_4$. Here, more than one mole of Mg_3N_2 reacts, which means more than these amounts of reactants and products are consumed and produced, respectively.

OWL: **3.3.2T: Tutorial Assignment**
3.3.2: Mastery Assignment

Molar mass is combined with stoichiometric factors in order to determine the mass of each species consumed or produced in the reaction. In these types of calculations, it is very important to keep track of the units for each value calculated.

For example, the mass of oxygen consumed by 7.8 g of C_4H_{10} is calculated as shown here. First, convert the mass of butane to an amount in units of moles.

$$7.8 \text{ g } C_4H_{10} \times \frac{1 \text{ mol } C_4H_{10}}{58.1 \text{ g}} = 0.13 \text{ mol } C_4H_{10}$$

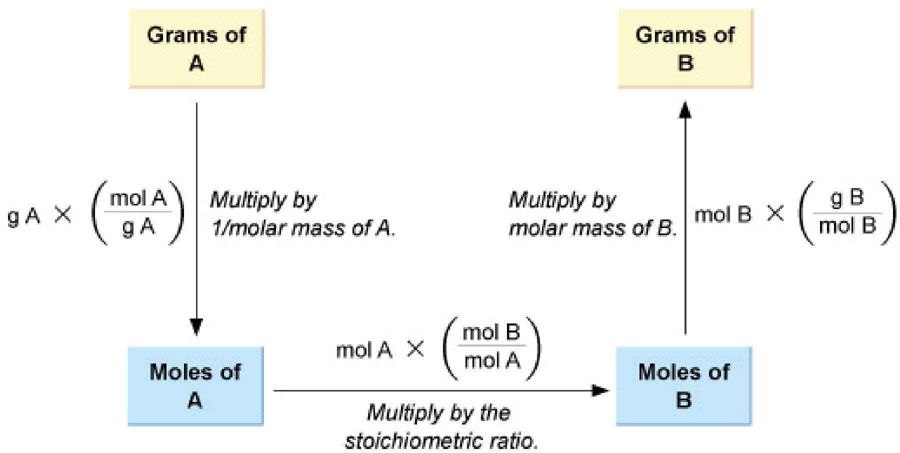
Next, use the stoichiometric factor from the balanced equation to convert moles of C_4H_{10} to moles of O_2.

$$0.13 \text{ mol } C_4H_{10} \times \frac{13 \text{ mol } O_2}{2 \text{ mol } C_4H_{10}} = 0.87 \text{ mol } O_2$$

Finally, use the molar mass of oxygen to calculate the mass of O_2 consumed by 7.8 g of C_4H_{10}.

$$0.87 \text{ mol } O_2 \times \frac{32.0 \text{ g}}{1 \text{ mol } O_2} = 28 \text{ g } O_2$$

In general, stoichiometric calculations are performed in the following order:



EXAMPLE PROBLEM: Use Reaction Stoichiometry to Calculate Amounts of Reactants and Products

Calcium carbonate reacts with hydrochloric acid to form calcium chloride, carbon dioxide, and water.

Unbalanced equation: $CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 + H_2O$

In one reaction, 54.6 g of CO_2 is produced. What amount (in mol) of HCl was consumed? What mass (in grams) of calcium chloride is produced?

SOLUTION:

You are asked to calculate the amount (in mol) of a reactant consumed and the mass (in grams) of a product formed in a reaction.

You are given the unbalanced equation and the mass of one of the products of the reaction.

Step 1. Write a balanced chemical equation.

$CaCO_3 + 2\ HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Step 2. Calculate moles of CO_2 produced.

$$54.6\text{ g }CO_2 \times \frac{1\text{ mol }CO_2}{44.01\text{ g}} = 1.24\text{ mol }CO_2$$

Step 3. Calculate moles of HCl consumed and moles of $CaCl_2$ produced using stoichiometric factors derived from the balanced chemical equation.

$$1.24\text{ mol }CO_2 \times \frac{2\text{ mol HCl}}{1\text{ mol }CO_2} = 2.48\text{ mol HCl}$$

$$1.24\text{ mol }CO_2 \times \frac{1\text{ mol }CaCl_2}{1\text{ mol }CO_2} = 1.24\text{ mol }CaCl_2$$

Step 4. Calculate mass (in grams) of $CaCl_2$ produced.

$$1.24\text{ mol }CaCl_2 \times \frac{111.0\text{ g}}{1\text{ mol }CaCl_2} = 138\text{ g }CaCl_2$$

Notice that it is possible to set up both calculations as a single step.

$$54.6\text{ g }CO_2\left(\frac{1\text{ mol }CO_2}{44.01\text{ g}}\right)\left(\frac{2\text{ mol HCl}}{1\text{ mol }CO_2}\right) = 2.48\text{ mol HCl}$$

$$54.6\text{ g }CO_2\left(\frac{1\text{ mol }CO_2}{44.01\text{ g}}\right)\left(\frac{1\text{ mol }CaCl_2}{1\text{ mol }CO_2}\right)\left(\frac{111.0\text{ g}}{1\text{ mol }CaCl_2}\right) = 138\text{ g }CaCl_2$$

Is your answer reasonable? More than one mole of CO_2 is produced in the reaction, so the amounts of HCl consumed and $CaCl_2$ produced are greater than one mole.

OWL: **3.3.3T: Tutorial Assignment**
3.3.3: Mastery Assignment

3.4 Stoichiometry and Limiting Reactants

Section Outline

3.4a Limiting Reactants
3.4b Percent Yield
Section Summary Assignment

The stoichiometric coefficients in a balanced equation are often not a reflection of the quantities combined in a typical laboratory experiment. For example, a common general chemistry lab experiment involves burning some magnesium metal in air to form solid magnesium oxide.

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

The balanced equation for this reaction shows that two moles of magnesium react with one mole of oxygen, but in the experiment the reactants are not combined in this ratio. Instead, much more oxygen is present than is needed to react completely with the magnesium. The study of stoichiometry includes not only reactions in which the reactants are completely consumed but also reactions that take place under nonstoichiometric conditions.

Opening Exploration 3.4 Limiting Reactants

When magnesium metal burns in air, the amount of product depends on how much magnesium is available, not on the amount of oxygen in the atmosphere.

3.4a Limiting Reactants

A balanced equation represents a situation in which a reaction proceeds to completion and all of the reactants are converted into products. Under nonstoichiometric conditions, the reaction continues only until one of the reactants is entirely consumed, and at this point the reaction stops. Under these conditions, it is important to know which reactant will be consumed first and how much product can be produced.

When a butane lighter is lit, for example, only a small amount of butane is released from the lighter and combined with a much larger amount of oxygen present in air.

$$2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$$

Under these conditions, the oxygen is considered an **excess reactant** because more is available than is required for reaction with butane. The butane is the **limiting reactant**, which means that it controls the amount of products produced in the reaction. When a nonstoichiometric reaction is complete, the limiting reactant is completely consumed and some amount of the excess reactant remains unreacted (Interactive Figure 3.4.1).

Interactive Figure 3.4.1 Determine How Much Zn Remains After Reaction with HCl

Reaction of Zn with HCl

Charles D. Winters

Consider a situation where 12 mol of butane is combined with 120 mol of oxygen. Under stoichiometric conditions, butane and oxygen react in a 2:13 mole ratio. The amount of oxygen needed to react with 10 mol of butane is

$$\text{mol } O_2 \text{ needed} = 12 \text{ mol } C_4H_{10} \times \frac{13 \text{ mol } O_2}{2 \text{ mol } C_4H_{10}} = 78 \text{ mol } O_2$$

Because more oxygen is available than is needed to react with all of the available butane (120 mol available versus 78 mol needed), oxygen is the excess reactant. The amount of butane that would be needed to react with 120 mol of O_2 is

$$\text{mol } C_4H_{10} \text{ needed} = 120 \text{ mol } O_2 \times \frac{2 \text{ mol } C_4H_{10}}{13 \text{ mol } O_2} = 18 \text{ mol } C_4H_{10}$$

Less butane is available than is needed to react with all of the oxygen available (12 mol available versus 18 mol needed), so butane is the limiting reactant. The amount of products that can be produced in the reaction is determined by the amount of limiting reactant present. For water,

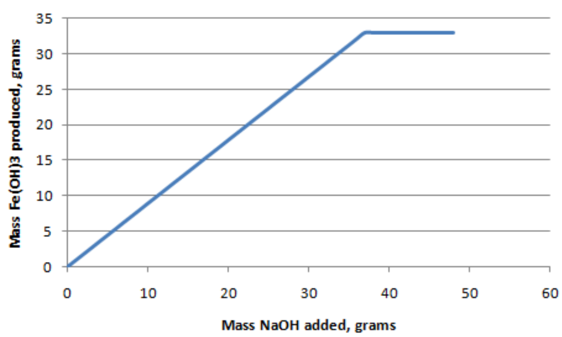
$$\text{mol } H_2O \text{ produced} = 12 \text{ mol } C_4H_{10} \times \frac{10 \text{ mol } H_2O}{2 \text{ mol } C_4H_{10}} = 60 \text{ mol } H_2O$$

The limiting reactant concept can be illustrated in the laboratory. For example, consider the reaction of iron(III) nitrate with sodium hydroxide.

$$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(s) + 3\ NaNO_3(aq)$$

If a student starts with 50.0 g of $FeCl_3$ and adds NaOH in 1-gram increments, the mass of $Fe(OH)_3$ produced in each experiment can be measured. Plotting mass of $Fe(OH)_3$ as a function of mass of NaOH added results in the graph shown in Interactive Figure 3.4.2.

Interactive Figure 3.4.2 Analyze Reaction Stoichiometry in the Laboratory



Mass of $Fe(OH)_3$ as a function of mass of NaOH

As expected, as more NaOH is added to the $FeCl_3$, increasing amounts of $Fe(OH)_3$ are produced. But when more than 37 g of NaOH is added, the mass of $Fe(OH)_3$ no longer changes. When 37 g (0.925 mol) of NaOH is added to 50 g (0.308 mol) of $FeCl_3$, the stoichiometric ratio is equal to 3:1 and all reactants are consumed completely. When more than 37 g of NaOH is added to 50 g $FeCl_3$, there is excess NaOH available ($FeCl_3$ is limiting) and no additional $Fe(OH)_3$ is produced.

EXAMPLE PROBLEM: Identify Limiting Reactants (Mole Ratio Method)

Identify the limiting reactant in the reaction of hydrogen and oxygen to form water if 61.0 g of O_2 and 8.40 g of H_2 are combined. Determine the amount (in grams) of excess reactant that remains after the reaction is complete.

SOLUTION:

You are asked to identify the limiting reactant and mass (in grams) of the excess reactant remaining after the reaction is complete.

You are given the mass of the reactants.

Step 1. Write a balanced chemical equation.

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

Step 2. Determine the limiting reactant by comparing the relative amounts of reactants available. Calculate the amount (in mol) of one of the reactants needed and compare that value to the amount available.

$$\text{Amount of } O_2 \text{ needed: } 8.40\text{ g } H_2\left(\frac{1\text{ mol } H_2}{2.016\text{ g}}\right)\left(\frac{1\text{ mol } O_2}{2\text{ mol } H_2}\right) = 2.08\text{ mol } O_2$$

$$\text{Amount of } O_2 \text{ available: } 61.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.00\text{ g}} = 1.91\text{ mol } O_2$$

More O_2 is needed (2.08 mol) than is available (1.91 mol), so O_2 is the limiting reactant.

Alternatively, the amounts of reactants available can be compared to the stoichiometric ratio in the balanced equation in order to determine the limiting reactant.

$$\frac{4.17\text{ mol } H_2}{1.91\text{ mol } O_2} = \frac{2.18\text{ mol } H_2}{1\text{ mol } O_2} > \frac{2\text{ mol } H_2}{1\text{ mol } O_2}$$

ratio of reactants (available) *ratio of reactants (balanced equation)*

[In equation above, align "ratio of reactants (available)" under " $\frac{2.18\text{ mol } H_2}{1\text{ mol } O_2}$ ***; align "ratio of reactants (balanced equation)" under*** $\frac{2\text{ mol } H_2}{1\text{ mol } O_2}$ ***]***

Here, the mole ratio of H_2 to O_2 available is greater than the mole ratio from the balanced chemical equation. Thus, hydrogen is in excess and oxygen is the limiting reactant.

Step 3. Use the amount of limiting reactant (O_2) available to calculate the amount of excess reactant (H_2) needed for complete reaction.

$$1.91\text{ mol } O_2\left(\frac{2\text{ mol } H_2}{1\text{ mol } O_2}\right)\left(\frac{2.016\text{ g}}{1\text{ mol } H_2}\right) = 7.70\text{ g } H_2$$

Step 4. Calculate the mass of excess reactant that remains when all the limiting reactant is consumed.

8.40 g H_2 available –7.70 g H_2 consumed = 0.70 g H_2 remains

Is your answer reasonable? More O_2 is needed than is available, so it is the limiting reactant. There is H_2 remaining after the reaction is complete because it is the excess reactant in the reaction.

OWL: **3.4.1T: Tutorial Assignment**
3.4.1: Mastery Assignment

An alternative method of performing stoichiometry calculations that involve limiting reactants is to simply calculate the maximum amount of product that could be produced from each reactant. (This method works particularly well for reactions with more than two reactants.) In this method, the reactant that produces the least amount of product is the limiting reactant.

EXAMPLE PROBLEM: Identify Limiting Reactants (Maximum Product Method)

Consider the reaction of gold with nitric acid and hydrochloric acid.

$$Au + 3\ HNO_3 + 4\ HCl \rightarrow HAuCl_4 + 3\ NO_2 + 3\ H_2O$$

Determine the limiting reactant in a mixture containing 125 g of each reactant and calculate the maximum mass (in grams) of $HAuCl_4$ that can be produced in the reaction.

SOLUTION:

You are asked to calculate the maximum mass of a product that can be formed in a reaction.

You are given the balanced equation and the mass of each reactant available.

In this case there are three reactants, so it is most efficient to calculate the maximum amount of product ($HAuCl_4$) that can be produced by each one. The reactant that produces the least amount of $HAuCl_4$ is the limiting reactant.

Step 1. Write the balanced chemical equation

$$Au + 3\ HNO_3 + 4\ HCl \rightarrow HAuCl_4 + 3\ NO_2 + 3\ H_2O$$

Step 2. Use the molar mass of each reactant and the stoichiometric factors derived from the balanced equation to calculate the amount of $HAuCl_4$ that can be produced by each reactant.

$$125\text{ g Au}\left(\frac{1\text{ mol Au}}{197.0\text{ g}}\right)\left(\frac{1\text{ mol HAuCl}_4}{1\text{ mol Au}}\right)\left(\frac{339.8\text{ g}}{1\text{ mol HAuCl}_4}\right) = 216\text{ g HAuCl}_4$$

$$125\text{ g HNO}_3\left(\frac{1\text{ mol HNO}_3}{63.01\text{ g}}\right)\left(\frac{1\text{ mol HAuCl}_4}{3\text{ mol HNO}_3}\right)\left(\frac{339.8\text{ g}}{1\text{ mol HAuCl}_4}\right) = 225\text{ g HAuCl}_4$$

$$125\text{ g HCl}\left(\frac{1\text{ mol HCl}}{36.46\text{ g}}\right)\left(\frac{1\text{ mol HAuCl}_4}{4\text{ mol HCl}}\right)\left(\frac{339.8\text{ g}}{1\text{ mol HAuCl}_4}\right) = 291\text{ g HAuCl}_4$$

Gold is the limiting reagent. The maximum amount of $HAuCl_4$ that can be produced is 216 g.

OWL: **3.4.2T: Tutorial Assignment**
3.4.2: Mastery Assignment

3.4b Percent Yield

When chemists perform experiments to make a compound, it is rare that the amount of materials produced is equal to that predicted based on stoichiometric calculations. The reasons for this are numerous. For example, side reactions may occur, using some of the reactants to make other, undesired products. Sometimes, the reaction does not go to completion and some reactants remain unreacted. In other cases, some of the product cannot be physically isolated (for example, because it is embedded in filter paper, it does not completely precipitate from solution or is lost via evaporation). Because laboratory results rarely match calculations exactly, we make a distinction between the predicted amount of product, called the **theoretical yield**, and the actual amount of product produced in the experiment, called the **experimental yield** (or *actual yield*).

When reporting the success of a chemical synthesis, the ratio of product mass obtained (the experimental yield) to the maximum mass that could have been produced based on the amounts of reactants used (the theoretical yield) is calculated. This ratio is reported in the form of a percent, called the **percent yield** of a reaction.

$$\text{percent yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\% \qquad (3.2)$$

All the calculations we have done to this point allow us to determine the theoretical yield for a reaction. An experimental yield value is never calculated but is instead measured in the chemical laboratory after a reaction is complete.

EXAMPLE PROBLEM: Calculate Percent Yield

Consider the reaction of $Pb(NO_3)_2$ with NaCl to form $PbCl_2$ and $NaNO_3$. If 24.2 g $Pb(NO_3)_2$ is reacted with excess NaCl and 17.3 g of $PbCl_2$ is ultimately isolated, what is the percent yield for the reaction?

SOLUTION:

You are asked to calculate the percent yield for a reaction.

You are given the mass of the limiting reactant and the experimental yield for one of the products in the reaction.

Step 1. Write a balanced chemical equation for the reaction.

$$Pb(NO_3)_2 + 2\ NaCl \rightarrow PbCl_2 + 2\ NaNO_3$$

Step 2. Use the amount of limiting reactant to calculate the theoretical yield of $PbCl_2$.

$$24.2\text{ g }Pb(NO_3)_2\left(\frac{1\text{ mol }Pb(NO_3)_2}{331.2\text{ g}}\right)\left(\frac{1\text{ mol }PbCl_2}{1\text{ mol }Pb(NO_3)_2}\right)\left(\frac{278.1\text{ g}}{1\text{ mol }PbCl_2}\right) = 20.3\text{ g }PbCl_2$$

Step 3. Use the experimental yield (17.3 g $PbCl_2$) and the theoretical yield (20.3 g $PbCl_2$) to calculate the percent yield for the reaction.

$$\frac{17.3\text{ g }PbCl_2}{20.3\text{ g }PbCl_2} \times 100\% = 85.2\%$$

Is your answer reasonable? The theoretical yield is greater than the experimental yield, so the percent yield is less than 100%.

3.5 Chemical Analysis

Section Outline

3.5a Determining a Chemical Formula
3.5b Analysis of a Mixture
Section Summary Assignment

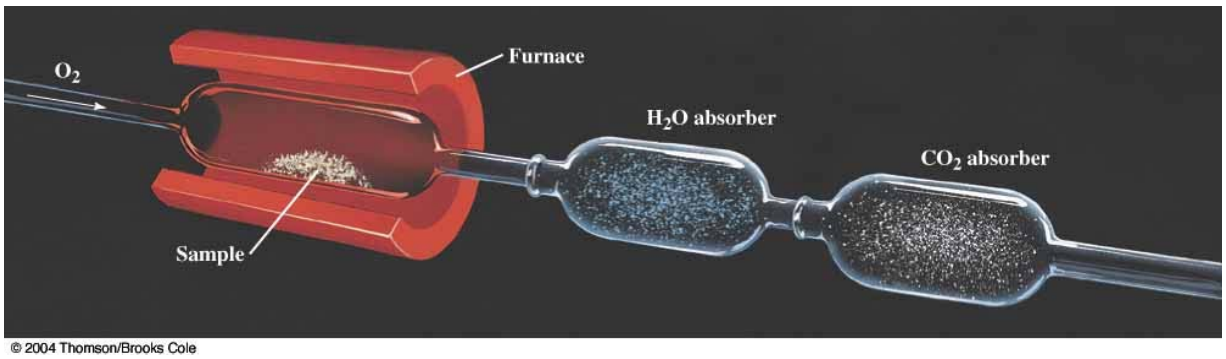
A large field of chemistry revolves around the need to analyze samples to determine the nature and amount of chemical species present. The chemical relationships presented so far in this chapter are essential to this kind of analysis. For example, the relationship between stoichiometry and chemical reactions is used to determine the amount of reactants that must be combined to synthesize a desired amount of product. The types of analysis done in a chemical laboratory, on the other hand, typically involve determining the chemical formula of an unknown compound, the composition of a mixture of compounds, or the purity of a chemical sample.

Opening Exploration 3.5 Element Analysis by Atomic Absorption Spectroscopy

3.5a Determining a Chemical Formula

The most common technique used to determine the chemical formula of compound that contains carbon and hydrogen is **combustion analysis**. In this technique, a weighed sample of the compound is burned in the presence of excess oxygen, converting all of the carbon in the sample to carbon dioxide and all of the hydrogen present in the sample to water (Interactive Figure 3.5.1). All the hydrogen in the sample ends up in the form of H_2O, which is absorbed in one chamber. All the carbon in the sample ends up in the form of CO_2, which is absorbed in the other chamber. The two absorption chambers are weighed before and after the reaction, and the chemical formula of the hydrocarbon can be determined.

Interactive Figure 3.5.1 CHO Combustion Analysis



Analysis of an organic compound by combustion

For the general hydrocarbon C_xH_y, the following reaction takes place during combustion analysis.

$$C_xH_y + \text{excess } O_2 \rightarrow x\, CO_2 + \frac{y}{2}\, H_2O$$

Using the stoichiometric relationships developed earlier in this chapter, the data produced from combustion analysis can be used to determine the empirical formula of a hydrocarbon, as shown in the following example.

EXAMPLE PROBLEM: Use Combustion Analysis to Determine Empirical and Molecular Formulas (Hydrocarbons)

When 1.827 g of a hydrocarbon, C_xH_y, was burned in a combustion analysis apparatus, 6.373 g of CO_2 and 0.7829 g of H_2O were produced. In a separate experiment, the molar mass of the compound was found to be 252.31 g/mol. Determine the empirical formula and molecular formula of the hydrocarbon.

SOLUTION:

You are asked to determine the empirical and molecular formulas of a compound.

You are given the mass of compound analyzed, data from a combustion analysis experiment (mass of H_2O and CO_2), and the molar mass of the compound.

Step 1. Use the mass of CO_2 and H_2O produced in the combustion analysis to calculate moles of H and moles of C present in the original sample. Note that there are 2 mol of H present in 1 mol of H_2O.

$$6.373 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.010 \text{ g}} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.1448 \text{ mol C}$$

$$0.7829 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g}} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.08692 \text{ mol H}$$

Step 2. Dividing each amount by the smallest value results in a whole-number ratio of elements.

$$\frac{0.1448 \text{ mol C}}{0.08692} = 1.666 \text{ mol C}$$

$$\frac{0.08692 \text{ mol H}}{0.08692} = 1.000 \text{ mol H} = 1 \text{ mol H}$$

In this case, the ratio is not made up of two integers. Writing the relative amounts as a ratio and rewriting the ratio as a fraction results in a whole-number ratio of elements and the correct empirical formula.

$$\frac{1.666 \text{ mol C}}{1 \text{ mol H}} = \frac{1\frac{2}{3} \text{ mol C}}{1 \text{ mol H}} = \frac{\frac{5}{3} \text{ mol C}}{1 \text{ mol H}} = \frac{5 \text{ mol C}}{3 \text{ mol H}}$$

The empirical formula of the hydrocarbon is C_5H_3.

Step 3. Compare the molar mass of the compound to the molar mass of the empirical formula to determine the molecular formula.

$$\frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{252.31 \text{ g/mol}}{63.079 \text{ g/mol}} = 4$$

The molecular formula is $(C_5H_3)_4$, or $C_{20}H_{12}$.

Is your answer reasonable? The molar mass of the compound must be a whole-number multiple of the molar mass of the empirical formula, which is true in this case.

Combustion analysis can also be performed on compounds containing carbon, hydrogen, and oxygen, as shown in the following example. In this case, the amount of oxygen must be determined from the mass of the original sample, not directly from the combustion data.

EXAMPLE PROBLEM: Use Combustion Analysis to Determine Empirical and Molecular Formulas (C, H, and O)

A 1.155-g sample of butyric acid, an organic compound containing carbon, hydrogen, and oxygen, is analyzed by combustion and 2.308 g of CO_2 and 0.9446 g of H_2O are produced. In a separate experiment, the molar mass is found to be 88.11 g/mol. Determine the empirical and molecular formulas of butyric acid.

SOLUTION:

You are asked to determine the empirical and molecular formulas of a compound.

You are given the mass of compound analyzed, data from a combustion analysis experiment (mass of H_2O and CO_2), and the molar mass of the compound.

When combustion analysis involves a compound containing carbon, hydrogen, and oxygen, the amount of oxygen in the compound cannot be determined directly from the amount of CO_2 or H_2O produced. Instead, the amounts of CO_2 and H_2O are used to calculate the mass of carbon and hydrogen in the original sample. The remaining mass of the original sample is oxygen.

Step 1. Use the mass of CO_2 and H_2O produced in the combustion analysis to calculate moles of C and moles of H present in the original sample. Note that there are 2 mol of H present in 1 mol of H_2O.

$$2.308 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.010 \text{ g}} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.05244 \text{ mol C}$$

$$0.9446 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g}} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.1049 \text{ mol H}$$

Step 2. Use the amount (in mol) of C and H to calculate the mass of C and H present in the original sample.

$$0.05244 \text{ mol C} \times \frac{12.011 \text{ g}}{1 \text{ mol C}} = 0.6299 \text{ g C}$$

$$0.1049 \text{ mol H} \times \frac{1.0079 \text{ g}}{1 \text{ mol H}} = 0.1057 \text{ g H}$$

Step 3. Subtract the amounts of C and H from the mass of the original sample to determine the mass of O present in the original sample. Use this value to calculate the moles of O present in the original sample.

$$1.155 \text{ g sample} - 0.6299 \text{ g C} - 0.1057 \text{ g H} = 0.419 \text{ g O}$$

$$0.419 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = 0.0262 \text{ mol O}$$

Step 4. Divide each amount by the smallest value to obtain a whole-number ratio of elements.

$$\frac{0.05244 \text{ mol C}}{0.0262} = 2.00 \text{ mol C} = 2 \text{ mol C}$$

$$\frac{0.1049 \text{ mol H}}{0.0262} = 4.00 \text{ mol H} = 4 \text{ mol H}$$

$$\frac{0.0262 \text{ mol O}}{0.0262} = 1.00 \text{ mol O} = 1 \text{ mol O}$$

The empirical formula of butyric acid is C_2H_4O.

Step 3. Compare the molar mass of the compound to the molar mass of the empirical formula to determine the molecular formula.

$$\frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}} = \frac{88.11 \text{ g/mol}}{44.05 \text{ g/mol}} = 2$$

The molecular formula is $(C_2H_4O)_2$, or $C_4H_8O_2$.

Is your answer reasonable? The molar mass of the compound must be a whole-number multiple of the molar mass of the empirical formula, which is true in this case.

3.5.2T: Tutorial Assignment

3.5.2: Mastery Assignment

The formula of other types of binary (two-element) compounds can be determined from experiments where the two elements react to form a single compound, as shown in the following example.

EXAMPLE PROBLEM: Determine the Chemical Formula of a Binary Compound

A 2.64-g sample of Cr is heated in the presence of excess oxygen. A metal oxide (Cr_xO_y) is formed with a mass of 3.86 g. Determine the empirical formula of the metal oxide.

SOLUTION:

You are asked to determine the empirical formula of a binary compound.

You are given the mass of a reactant and the mass of the binary compound formed in the reaction.

Step 1. Use the mass of the metal oxide and the mass of the metal to determine the amount of oxygen present in the metal oxide sample.

$3.86 \text{ g } Cr_xO_y - 2.64 \text{ g Cr} = 1.22 \text{ g O}$

Step 2. Use the mass of Cr and mass of O to determine moles of each element present in the compound.

$$2.64 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{52.00 \text{ g}} = 0.0508 \text{ mol Cr}$$

$$1.22 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g}} = 0.0763 \text{ mol O}$$

Step 3. To determine the simplest whole-number ratio of elements, divide each by the smallest value.

$$\frac{0.0508 \text{ mol Cr}}{0.0508} = 1.00 \text{ mol Cr} = 1 \text{ mol Cr}$$

$$\frac{0.0763 \text{ mol O}}{0.0508} = 1.50 \text{ mol O}$$

In this case, the ratio is not made up of two integers. Writing the relative amounts as a ratio and rewriting the ratio as a fraction results in a whole-number ratio of elements and the correct empirical formula.

$$\frac{1.50 \text{ mol O}}{1 \text{ mol Cr}} = \frac{\frac{3}{2} \text{ mol O}}{1 \text{ mol Cr}} = \frac{3 \text{ mol O}}{2 \text{ mol Cr}}$$

The empirical formula is Cr_2O_3.

3.5b Analysis of a Mixture

The analysis of a mixture is a common process performed in the chemical laboratory. For example, an environmental sample containing a mixture of compounds can be analyzed for a single compound by first chemically isolating the compound of interest, followed by chemical analysis to determine the amount of compound in the sample. This type of analysis uses the stoichiometric methods developed in this chapter to determine the purity of a chemical sample or the chemical makeup of the sample, as shown in the following example.

EXAMPLE PROBLEM: Use Stoichiometric Methods to Analyze a Mixture

A soil sample is analyzed to determine the iron content. First, the iron is isolated as $Fe(NO_3)_2$. This compound is then reacted with $KMnO_4$.

$$5\ Fe(NO_3)_2 + KMnO_4 + 8\ HNO_3 \rightarrow 5\ Fe(NO_3)_3 + Mn(NO_3)_2 + 4\ H_2O + KNO_3$$

In one experiment, a 12.2-g sample of soil required 1.73 g of $KMnO_4$ to react completely with the $Fe(NO_3)_2$. Determine the percent (by mass) of iron present in the soil sample.

SOLUTION:

You are asked to calculate the mass percent of iron in a soil sample.

You are given a balanced equation, the mass of soild analyzed, and the mass of one of the reactants used to analyze the sample.

Step 1. Use the mass of $KMnO_4$ to determine the amount of $Fe(NO_3)_2$ present in the soil sample.

$$1.73\text{ g } KMnO_4 \left(\frac{1\text{ mol } KMnO_4}{158.0\text{ g}}\right)\left(\frac{5\text{ mol } Fe(NO_3)_2}{1\text{ mol } KMnO_4}\right) = 0.0547\text{ mol } Fe(NO_3)_2$$

Step 2. Use the stoichiometry of the chemical formula of $Fe(NO_3)_2$ and the molar mass of Fe to calculate the mass of Fe present in the soil sample.

$$0.0547\text{ mol } Fe(NO_3)_2 \left(\frac{1\text{ mol Fe}}{1\text{ mol } Fe(NO_3)_2}\right)\left(\frac{55.85\text{ g}}{1\text{ mol Fe}}\right) = 3.06\text{ g Fe}$$

Step 3. Use the mass of Fe in the soil sample and the mass of the soil sample to calculate the percent by mass of Fe in the soil.

$$\frac{3.06\text{ g Fe in soil sample}}{12.2\text{ g soil sample}} \times 100\% = 25.1\%\text{ Fe}$$

Is your answer reasonable? Because the soil sample is not pure, the mass of iron in the soil is less than the mass of the soil sample and the weight percent of iron in the sample is less than 100%.

OWL: **3.5.4T: Tutorial Assignment**
3.5.4: Mastery Assignment

Key Concepts

3.1 The Mole and Molar Mass

- Avogadro's number (6.022×10^{23}) is the number of particles in one mole of a substance (3.1a).
- Molar mass is the mass in grams of one mole of particles of a substance (3.1b).
- Molar mass is used to convert between moles and grams of a substance (3.1b).

3.2 Stoichiometry and Compound Formulas

- Stoichiometry is the study of the relationship between relative amounts of substances (3.2a).
- Percent composition of an element in a compound is the mass of an element in exactly 100 g of a compound (3.2b).
- The empirical formula of a compound is the simplest whole-number ratio of elements present in the compound (3.2c).
- The molecular formula of a compound is a whole-number multiple of the empirical compound formula and shows the number of atoms of each element in one molecule of a compound (3.2d).
- Hydrated ionic compounds have water trapped in the crystal lattice (3.2e).
- The formula of a hydrated compound can be determined by heating the compound to drive off waters of hydration (3.2e).

3.3 Stoichiometry and Chemical Reactions

- Chemical reactions are represented by chemical equations, in which the reacting species (reactants) are shown to the left of a reaction arrow and the species produced during the reaction (products) appear to the right of the reaction arrow (3.3a).
- Stoichiometric coefficients are numbers that appear to the left of a compound formula in a chemical equation (3.3a).
- A balanced equation, which has equal numbers of atoms of each element present in the reaction on both sides of the reaction arrow, reflects the Law of Conservation of Matter, which states that matter is neither created nor destroyed during a chemical reaction (3.3a).
- The stoichiometric coefficients in a balanced chemical equation can be used to create a stoichiometric factor (or stoichiometric ratio) that relates the amount of one species to another (3.3c).
- Using stoichiometric factors and molar mass, it is possible to determine the amounts of reactants or products consumed or produced in a chemical reaction (3.3c).

3.4 Stoichiometry and Limiting Reactants

- The limiting reactant is the reactant in a chemical reaction that limits or controls the amount of product that can be produced (3.4a).
- Excess reactants are present in amounts that exceed the amount required for reaction (3.4a).
- Percent yield is a ratio of the experimental yield (the amount of material produced in the reaction) to the theoretical yield (the predicted amount of product) (3.4b).

3.5 Chemical Analysis

- Combustion analysis is an analytical method used to determine the carbon and hydrogen content, by mass, in a compound (3.5a).

Key Equations

$$\%\ \text{element} = \frac{(\text{number of atoms of element in formula})(\text{molar mass of element})}{\text{mass of 1 mol of compound}} \times 100\% \quad (3.1)$$

$$\text{percent yield} = \frac{\text{experimental yield}}{\text{theoretical yield}} \times 100\% \quad (3.2)$$

Key Terms

3.1 The Mole and Molar Mass

mole (mol)
Avogadro's number (N_A)
molar mass
formula weight
molecular weight

3.2 Stoichiometry and Compound Formulas

stoichiometry
percent composition
empirical formula
molecular formula

3.3 Stoichiometry and Chemical Reactions

chemical equations
reactants
products
stoichiometric coefficient
Law of Conservation of Matter
balanced equation
stoichiometric factor

3.4 Stoichiometry and Limiting Reactants

excess reactant
limiting reactant
theoretical yield
experimental yield
percent yield

3.5 Chemical Analysis

combustion analysis