Beginning Chemistry
A Guide to the Fundamentals

An Honors Thesis (HONRS 499)

by

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Preface

I have been a tutor of general chemistry for one and a half years. During that time, I have noticed patterns in the types of questions that students most commonly ask. I have also learned which types of problems most often give students difficulty. For these reasons, I have decided to assemble a tutorial designed to specifically address those questions and those problems. These are the topics that I feel are most essential for the student who is having trouble in a beginning chemistry course. This tutorial focuses on the skills and strategies involved in problem solving. I have attempted to help the student visualize the steps that lead to the solutions to certain types of chemistry problems. I have also emphasized the similarities between problems that might otherwise seem very different. In this way, I hope that the student will eventually approach these types of problems with confidence and no longer wonder how to begin solving them.

Finally, I must emphasize that this booklet should not be used in place of the student's regular class materials. I have not written this tutorial as a comprehensive guide to all aspects of general chemistry. Rather, my hope is that it will serve as a helpful supplement to the student's regular textbook and class lectures. When it is used in this way, I believe that this booklet will answer many of the student's questions about beginning chemistry and explain difficult concepts more clearly.

Scott D. Thole

Contents

I. Significant Figures

Constants
Experimental Measurements
Significant Figures
Calculations with Significant Figures

Before a student begins to work problems in Chemistry, he must first understand the nature of the numbers involved in the related calculations. There are two basic types of numbers that the student must be able to identify:

1. Constants
2. Experimental Measurements

The fundamental difference between these two types of numbers is the degree of certainty with which we employ their values in a calculation.

Constants

A **constant** is a number that is considered to be *exact*. We therefore assume that *every* digit in a **constant** is known with certainty. This category of numbers includes both physical constants and conversion factors. The following values are examples of physical constants:

NAME	**SYMBOL**	**VALUE**
Avogadro's number	N	6.02×10^{23} mol^{-1}
Gas constant	R	8.314 J K^{-1} mol^{-1}
Planck's constant	h	6.626×10^{-34} J s

These accuracy of these values is considered to be exact when they are used in calculations. For example, the Gas constant has a value of 8.314 J K^{-1} mol^{-1}. Because this number is accepted as a constant, we assume that *every* digit is known with certainty. As a result, the '8', '3', '1', and '4' are all considered exact. The number is accepted as '8.314', and its accuracy is not questioned in calculations.

Conversion factors include numbers that exist by definition. The following values are examples of defined numbers:

1 week = 7 days
1 yard = 3 feet
1 dime = 10 cents
1 inch = 2.54 centimeters

1 kilogram = 2.2046 pounds

These relationships are known with certainty and the values are therefore considered exact when used in calculations. For example, the number of days in a week is not questioned because a week contains seven days by definition. Thus, there are *exactly* seven days in one week. Similarly, because a yard contains three feet by definition, there are *exactly* 3 feet in one yard. Because a dime has been defined as ten cents, there are *exactly* ten cents in one dime. Furthermore, the inch has recently been defined as 2.54 centimeters. Thus, there are *exactly* 2.54 centimeters in one inch. In other words, because these numbers exist by definition, their accuracy cannot be questioned.

The relationship between kilograms and pounds, however, requires special attention. This example is different from the others in that the kilogram was not defined *in terms of* pounds. Rather, the relationship between the kilogram and the pound is the result of the independent definitions of the two units. Nevertheless, an exact relationship between these two units results from their independent definitions. However, the number of pounds in one kilogram may be indicated more accurately than in the above example. Nevertheless, although the number '2.2046' is only an estimate, the relationship between these units is still the result of the independent definitions of the kilogram and the pound. The number is therefore assumed to be exact when used in calculations.

Measurements

A **experimental measurement** is not known with certainty to be exact. Furthermore, the value of a measurement is not the result of definition. Therefore, unlike constants, measurements are not assumed to be exact. A measurement is only considered accurate within a given range of possible error. Thus, every measurement includes one estimated digit. The range of possible error for a given measurement and the position of the estimated digit depend on the accuracy of the measuring device. An experimental measurement is therefore a value that includes an estimated digit which is not known with certainty. The following example is one of many types of measurements:

length = 15.8 mm

This value is a typical measurement that would be obtained using a ruler with millimeter graduations. In this example, the '8' is an estimated digit since there are no graduations on the ruler smaller than 1 mm. Another person might have estimated the length to be 15.7 mm, and yet another person might have estimated the value to be 15.9 mm. Therefore, the measurement is considered accurate only within the limits of error caused by the estimation of the last digit. (Note that in all cases the '15' is known with certainty.) Thus, unlike a constant, a measurement is an experimental value that always contains an estimated digit that is uncertain.

Significant Figures

Significant figures are defined as all the digits in an **experimental measurement** that are known with certainty, plus one digit that has been estimated, and not including zeros which serve only to place the decimal (i.e. zeros that would disappear in scientific notation). For any measured value, the number of digits fitting this description is the number of significant figures in that measurement. Note that the concept of significant figures applies only to measurements. It is incorrect to describe **constants** in terms of significant figures. Recall the previous example of an experimental measurement:

length = 15.8 mm

In this example, the '1' and the '5' are known with certainty, and the '8' has been estimated. Thus, this measurement contains three significant figures. Study the following examples very carefully. In each measurement, the estimated digit is printed in **bold** face, and any zeros which serve only to place the decimal are printed in *italics*:

MEASURED VALUE	NUMBER OF SIGNIFICANT FIGURES
2_00_	1
2_0_	1
2	1
_0._**2**	1
_0.0_**2**	1
_0.00_**2**	1
_0.0000_**2**	1
12**0**	2
1**2**	2
1.**2**	2
2.**0**	2
_0._2**0**	2
_0.0_2**0**	2
_0.000_2**0**	2
12.**2**	3
2.0**0**	3
2.0**1**	3
2.1**0**	3
_0._12**2**	3
_0.0_12**0**	3
_0.00_12**0**	3
200.**1**	4
20.0**1**	4
20.1**0**	4
20.0**0**	4
2.00**0**	4
_0._200**0**	4
_0.0_200**0**	4
2000.**0**	5
201.0**1**	5
21.10**0**	5
20.00**1**	5
2.000**0**	5
_0._2001**0**	5
_0._2000**0**	5

EXAMPLE 1:

Determine the number of significant figures in each of the following measurements:

a) 2.5 cm	f) 999.90 ml
b) 0.0333 m	g) 15.01 yards
c) 1.75 x 10-3 kg	h) 50.00 feet
d) 0.00175 kg	i) 600 m
e) 40.0 in	j) 600.0 m

SOLUTIONS:

a) 2	f) 5
b) 3	g) 4
c) 3	h) 4
d) 3	i) 1
e) 3	j) 4

Calculations with Significant Figures

When performing calculations that involve measured numbers, it is important to realize that we cannot improve accuracy through mathematical manipulations of measurements. The limitations of our measuring device restrict our level of accuracy even when the measurement becomes involved in calculations. Therefore, the number of significant figures possible in the answer to a particular calculation depends upon the number of significant figures present in the measurements that led to that answer. In order to determine the number of significant figures possible in the answer to a calculation involving measured values, we must follow two rules:

1) ADDITION/SUBTRACTION: The answer must have the same number of *decimal places* as that *measurement* with the *fewest* number of decimal places.
2) MULTIPLICATION/DIVISION: The answer must have the same number of *significant figures* as that *measurement* with the *fewest* number of significant figures

NOTE: *Constants* do not affect the number of significant figures in the answer.

EXAMPLE 2:

Perform the following calculations and express the answer to the correct number of significant figures:

a) 3.333 + 8.1 + 2.50 + 16.66
b) 2.625 x 100 / 1.05
c) (34.16 - 32.50) x 4.175

SOLUTIONS:

a) 30.6 (one decimal place)
b) 300 (one significant figure)
c) 6.93 (three significant figures)

II. Dimensional Analysis

Dimensional analysis, also called the **unit factor conversion method**, is a technique used to convert the units of a particular measurement without changing the actual physical quantity it represents. For example, dimensional analysis can be used to convert a person's weight in pounds to his weight in kilograms. Although the value will be expressed in different units, the weight of the person does not change.

Suppose a woman weighs 125 pounds, and she wishes to know her weight in kilograms. In order to perform this conversion, we must first refer to a table of conversion factors to find out the number of kilograms in a pound. Such a table will provide the following information:

1 pound = 0.45359 kilogram

Thus, a table of conversion factors provides the *numbers* for the unit conversion factor. A **unit conversion factor** is a ratio of *equal quantities* expressed in *different units*. It is always possible to create at least two unit conversion factors from the numbers given in the table. In this example, the relationship between the pound and the kilogram shown above can be used to create the following unit conversion factors:

$$\frac{\text{1 pound}}{\text{0.45359 kg}} \qquad \frac{\text{0.45359 kg}}{\text{1 pound}}$$

Note that these two unit conversion factors have the same *numbers*, but differ in their *orientation*. The orientation that must be used depends on the units you wish to obtain after the conversion. The next step is to multiply the starting value by the appropriate unit conversion factor, being careful to select the orientation that will cause the starting units to cancel out:

$$125 \text{ pounds} \times \frac{0.45359 \text{ kg}}{1 \text{ pound}} = 56.7 \text{ kg}$$

Because the starting value has 'pounds' in the numerator, we must orient the unit conversion factor with 'pounds' in the denominator. This causes 'pounds' to cancel out and 'kilograms' to remain.

It is important to realize that, although the number has decreased, the woman has not lost weight as a result of the conversion process. Remember that the unit conversion factor is a ratio of *equal quantities*, and that any ratio in which the numerator equals the denominator is equivalent to '1'. Thus, the unit conversion factor is equal to '1', just as the following ratios are equal to one:

$$\frac{2}{2} = \frac{3}{3} = \frac{4}{4} = \frac{x}{x} = \frac{a + b}{a + b} = \frac{0.45359 \text{ kg}}{1 \text{ pound}} = 1$$

Because '1 pound' equals '0.45359 kg', the unit conversion factor used in the above example must also equal '1'. That means that our starting value was multiplied by the equivalent of '1'. The result therefore represents the *same physical quantity* of weight in *different units*.

EXAMPLE 1:

Answer the following questions using the principles of dimensional analysis and the appropriate unit conversion factor(s):

a) How many quarters equal $6.75 ?
b) How many minutes are in one year?
c) How many wheels are required to build a dozen tricycles?

SOLUTIONS:

a)

$$6.75 \text{ dollars} \times \frac{4 \text{ quarters}}{1 \text{ dollar}} = 27 \text{ quarters}$$

b)

$$1 \text{ year} \times \frac{365 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{60 \text{ minutes}}{1 \text{ hour}} = 525{,}600 \text{ minutes}$$

c)

$$1 \text{ dozen tricycles} \times \frac{12}{1 \text{ dozen}} \times \frac{3 \text{ wheels}}{1 \text{ tricycle}} = \mathbf{36\ wheels}$$

EXAMPLE 2:

Dimensional analysis is also very helpful in making conversions between units *within* the SI system. Use the appropriate unit conversion factor(s) to perform the following conversions:

a) 50 m to mm
b) 50 mm to m
c) 900 g to kg
d) 25 cm to km
e) 4.5 nm to m
f) 6.6 kg to mg
g) 2.0 m^3 to cm^3
h) 1500 cm^3 to L

SOLUTIONS:

a) NOTE: When we convert to ***smaller units***, we should expect a ***larger number***.

$$50 \text{ m} \times \frac{1000 \text{ mm}}{1 \text{ m}} = \mathbf{50{,}000\ mm}$$

b) NOTE: When we convert to ***larger units***, we should expect a ***smaller number***.

$$50 \text{ mm} \times \frac{1 \text{ m}}{1000 \text{ mm}} = \mathbf{0.05\ m}$$

c)

$$900 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = \mathbf{0.9\ kg}$$

d)

$$25 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ km}}{1000 \text{ m}} = \mathbf{2.5 \times 10^{-4}\ km}$$

e)

$$4.5 \text{ nm} \times \frac{1 \text{ m}}{10^9 \text{ nm}} = \mathbf{4.5 \times 10^{-9}\ m}$$

f)

$$6.6 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1000 \text{ mg}}{1 \text{ g}} = \mathbf{6.6 \times 10^{6}\ mg}$$

g) NOTE: Pay special attention to the unit conversion factor used in this example.

$$2.0 \text{ m}^3 \times \frac{100^3 \text{ cm}^3}{1 \text{ m}^3} = \mathbf{2.0 \times 10^{6}\ cm^3}$$

h)

$$1500 \text{ cm}^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} \times \frac{1 \text{ L}}{1000 \text{ mL}} = \mathbf{1.5\ L}$$

EXAMPLE 3:

Use dimensional analysis and a table of conversion factors to perform the following conversions between the English system and the SI system:

a) 100 yards to m
b) 18 gallons to mL
c) 2.4 tons to g
d) 1 inch to nm

SOLUTIONS:

a)

$$100 \text{ yards} \times \frac{1 \text{ m}}{1.0936 \text{ yards}} = 91.4 \text{ m}$$

b)

$$18 \text{ gallons} \times \frac{3.7854 \text{ L}}{1 \text{ gallon}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 6.8 \times 10^4 \text{ ml}$$

c)

$$2.4 \text{ tons} \times \frac{907.185 \text{ kg}}{1 \text{ ton}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 2.2 \times 10^6 \text{ g}$$

d)

$$1 \text{ inch} \times \frac{2.54 \text{ cm}}{1 \text{ inch}} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 2.54 \times 10^7 \text{ nm}$$

III. The Mole

The concept of the **mole** is often confusing for beginning chemistry students. I have found that it is often helpful to compare the mole to the concept of the **dozen**. We are all familiar with the idea of a dozen. A dozen represents twelve. If someone asks for a dozen eggs, we know that he wants twelve eggs. The concept of the mole is very similar to the idea of a dozen. The 'mole' is a term that chemists use to represent a particular number just as the term 'dozen' represents twelve. However, the number represented by a 'mole' is much larger than twelve. A 'mole' represents 6.02×10^{23}. Thus, if a chemist asks for a *mole* of eggs, we know that he wants 6.02×10^{23} eggs. Similarly, if the chemist asks for a mole of hydrogen atoms, we know that he wants 6.02×10^{23} atoms of hydrogen. And if he demands a mole of water, we must give him 6.02×10^{23} molecules of H_2O. In other words, a 'mole' is like a 'dozen', except a 'mole' does not mean twelve. A 'mole' means 6.02×10^{23}.

EXAMPLE 1:
A chemist claims to have found 2.0 moles of gold. How many gold atoms does he now posses?

SOLUTION:

$$2.0 \text{ moles} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}} = 1.2 \times 10^{24} \text{ atoms}$$

Thus, the chemist has found 1.2×10^{24} atoms of gold.

EXAMPLE 2:
Your lab instructor has produced 100 molecules of carbon dioxide during a combustion. How many moles of CO_2 has he produced?

SOLUTION:

$$100 \text{ molecules} \times \frac{1 \text{ mole}}{6.02 \times 10^{23} \text{ molecules}} = 1.66 \times 10^{-22} \text{ moles}$$

Thus, your instructor has produced 1.66×10^{-22} moles of CO_2.

The number 6.02 x 10^{23} has a special name. It is referred to as Avogadro's number and is often abbreviated as N_A. The number 6.02 x 10^{23} was not selected arbitrarily to be represented by the 'mole'. Chemists decided to define the mole as the number of carbon atoms found in twelve grams of carbon-12. If a person has a mole of water molecules, he has the same number of water molecules as there are carbon atoms in twelve grams of carbon-12. Based on this definition of the mole, we can say that a mole of carbon-12 weighs exactly 12 grams. This brings us to the concept of **molar weight**. The molar weight of an element is simply its atomic weight in grams. The atomic weights of each of the elements are found on a periodic table. For example, the atomic weight of nitrogen is 14. Thus, the molar weight of nitrogen is 14 ***grams***. In other words, one mole of nitrogen atoms weighs 14 grams. **Molecular weight** is the sum of the atomic weights of all the atoms present in a given molecule. The molar weight of a *molecule* then is its molecular weight in grams. However, chemists usually refer to the 'molar weight of a molecule' as its 'molecular weight'. In other words, chemists generally use the term 'molecular weight' to refer to the weight of one mole of the molecules. For example, the *molar weight* of H_2O is 18 grams. It takes 6.02 x 10^{23} molecules of H_2O to equal 18 grams. However, chemists will say that the 'molecular weight' of H_2O is 18 grams even though this is not the weight of one molecule of water. Thus, the term 'molecular weight' will be used to represent the molar weight of the molecule.

It will often be necessary for students to use these concepts in order to convert between moles, grams, and number of particles of a particular substance when working problems. The following diagram is often helpful in directing the necessary conversions between these units:

GRAMS ⟷ **MOLES** ⟷ NUMBER OF PARTICLES

As shown by the diagram, it is possible to convert directly between weight and 'moles' and between 'moles' and 'number of particles', but it is most convenient to pass through 'moles' when converting between weight and 'number of particles'.

EXAMPLE 1:

A weighing dish contains 1.00 g of $MgSO_4$. Find the number of moles of $MgSO_4$ present in the weighing dish.

SOLUTION:

$$1.00 \text{ g } MgSO_4 \times \frac{1 \text{ mol}}{120 \text{ g}} = 8.33 \times 10^{-3} \text{ mol } MgSO_4$$

Thus, the dish contains 8.33×10^{-3} moles of $MgSO_4$.

EXAMPLE 2:

A reaction is expected to produce 2.20 moles of NH_3. What weight of NH_3 should be produced during this reaction?

SOLUTION:

$$2.20 \text{ mol } NH_3 \times \frac{17 \text{ g}}{\text{mol}} = 37.4 \text{ g } NH_3$$

Thus, the reaction should produce 37.4 grams of NH_3.

EXAMPLE 3:

80.0 milligrams of potassium phosphate were used to prepare a buffer. How many molecules of the compound were used?

SOLUTION:

$$80.0 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 0.0800 \text{ g}$$

$$0.0800 \text{ g } K_3PO_4 \times \frac{1 \text{ mol}}{212 \text{ g}} = 3.77 \times 10^{-4} \text{ mol } K_3PO_4$$

$$3.77 \times 10^{-4} \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 2.27 \times 10^{20} \text{ molecules}$$

Thus, 2.27×10^{20} molecules of K_3PO_4 were used in the buffer.

IV. Percent Composition

Empirical Formula
Molecular Formula

When a student has learned to work problems that involve **moles and molecular weight,** he is ready encounter the concept of **percent composition.** Percent composition is generally expressed as percent by weight and is used to describe the relative amount of each type of atom in a particular molecule. For example, the percent by weight of carbon in a molecule of carbon dioxide indicates the percent of the total weight of a CO_2 molecule that is attributed to the carbon atom. Similarly, because all molecules of CO_2 have the same percent composition, we can say that the percent of carbon in a tank filled with CO_2 is the percent of the total weight of CO_2 in the tank that is due to all of the carbon atoms. A similar analysis could be performed for the oxygen portion of the CO_2 molecule. The percent composition of a molecule is determined by dividing the total weight of each type of atom in the molecule by the total weight of the molecule. In other words, the percent composition of a molecule is calculated by dividing the total weight of each element in the molecule by the **molecular weight.** Study each of the following examples very carefully.

EXAMPLE 1:
Determine the percent composition of the carbon dioxide molecule.

SOLUTION:

carbon dioxide: CO_2
molecular weight CO_2: 44 g (1 mole CO_2 = 44 g)

molar weight of C: 12 g — number of C atoms per molecule: 1
molar weight of O: 16 g — number of O atoms per molecule: 2

$$\frac{12\text{ g}}{\text{mole C atoms}} \times \frac{1\text{ C atom}}{\text{molecule } CO_2} = 12\text{ g carbon/mole } CO_2$$

$$\frac{12\text{ g carbon}}{\text{mole } CO_2} \times \frac{\text{mole } CO_2}{44\text{ g}} \times 100\% = 27\%\text{ carbon}$$

$$\frac{16\ g}{\text{mole O atoms}} \times \frac{2\ \text{O atoms}}{\text{molecule } CO_2} = 32\ \text{g oxygen/mole } CO_2$$

$$\frac{32\ \text{g oxygen}}{\text{mole } CO_2} \times \frac{\text{mole } CO_2}{44\ g} \times 100\% = 73\%\ \text{oxygen}$$

Percent composition of CO_2: 27% C, 73% O.

EXAMPLE 2:
Determine the percent of sulfuric acid that is sulfur.

SOLUTION:

sulfuric acid: H_2SO_4
molecular weight H_2SO_4: 98 g (1 mole H_2SO_4 = 98 g)

molar weight of S: 32 g number of S atoms per molecule: 1

$$\frac{32\ g}{\text{mole S atoms}} \times \frac{1\ \text{S atom}}{\text{molecule } H_2SO_4} = 32\ \text{g sulfur/mole } H_2SO_4$$

$$\frac{32\ \text{g sulfur}}{\text{mole } H_2SO_4} \times \frac{\text{mole } H_2SO_4}{98\ g} \times 100\% = 33\%\ \text{sulfur}$$

Sulfuric acid is 33% sulfur by weight.

Empirical Formula

The percent composition of a molecule is often used to determine its **empirical formula.** An empirical formula indicates only the **relative number** of each type of atom in a molecule. The empirical formula does *not* indicate the **actual number** of each type of atom in the molecule. The following example illustrates how one can determine the empirical formula of a molecule from its percent composition.

EXAMPLE 3:
A chemist has determined that a particular unknown compound is 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen. From this percent composition, determine the empirical formula of the compound.

SOLUTION:

The following table will be helpful in working this type of problem.

atom	percent	100 g sample	moles	ratio
C	40.0%			
H	6.7%			
O	53.3%			

The first step is to predict the number of grams of each type of atom that would be present in a 100 g sample of the unknown compound. By selecting a theoretical sample that weighs 100 g, we can greatly simplify the necessary calculations in this step. Because the unknown compound is 40.0% carbon by weight, a 100 g sample would contain 40.0 g carbon. Similarly, because the compound is 6.7% hydrogen and 53.3% oxygen, a 100 g sample would contain 6.7 g hydrogen and 53.3 g oxygen. These values can now be placed in the table as follows:

atom	percent	100 g sample	moles	ratio
C	40.0%	40.0 g		
H	6.7%	6.7 g		
O	53.3%	53.3 g		

The next step is to convert each weight to moles by dividing by the molar weight of each atom respectively. The number of moles of each atom can then be placed in the next column of the table.

atom	percent	100 g sample	moles	ratio
C	40.0%	40.0 g	3.3	
H	6.7%	6.7 g	6.7	
O	53.3%	53.3 g	3.3	

The final step is to determine the whole number ratio of each type of atom in the sample by dividing each number of moles by the lowest number of moles in the column. The lowest number of moles in this example is 3.3. When each number of moles is divided by 3.3, we obtain the following whole number ratio of atoms as shown in the completed table below.

atom	percent	100 g sample	moles	ratio
C	40.0%	40.0 g	3.3	1
H	6.7%	6.7 g	6.7	2
O	53.3%	53.3 g	3.3	1

The whole number ratios obtained in the last column of the completed table are the subscripts of the empirical formula. These numbers indicate the relative number of each type of atom in a molecule of the unknown compound. We can therefore conclude that the empirical formula for this compound is $C_1H_2O_1$ or $\mathbf{CH_2O}$.

Molecular Formula

The empirical formula of an unknown compound can now be used to determine its **molecular formula**. Unlike an empirical formula, which indicates only the *relative number* of each type of atom in a molecule, the **molecular formula** indicates the *actual number* of each type of atom in the molecule. In order to determine the molecular formula of a compound from its empirical formula, one must also know the **molecular weight** of the compound. The molecular formula is determined by comparing the theoretical weight of the empirical formula to the known molecular weight. If the known molecular weight equals the theoretical weight of the empirical formula, the molecular formula equals the empirical formula. However, if the known molecular weight is higher than the theoretical weight of the empirical formula, the molecular formula will be a multiple of the empirical formula. For example, if the known molecular weight of a compound is twice the theoretical weight of the empirical formula, the molecular formula will be the empirical formula times two. Note that the known molecular weight will always be **greater than** or **equal to** the theoretical weight of the empirical formula. The known molecular weight is *never* less than the theoretical weight of the empirical formula. Thus, the molecular formula is always a whole number multiple of the empirical formula.

EXAMPLE 4:

The empirical formula of the unknown compound in the previous example was found to be CH_2O. a) Find the **molecular formula** of this compound if its molecular weight is known to be **30 g/mol**. b) What would be the **molecular formula** of the compound if its molecular weight were determined to be **180 g/mol**?

SOLUTIONS:

a) molecular weight: 30 g/mol
empirical formula: CH_2O
CH_2O theoretical weight: 30 g/mol

NOTE: theoretical weight = molecular weight

Thus, the molecular formula is the same as the empirical formula.

molecular formula: CH_2O (formaldehyde)

b) molecular weight: 180 g/mol
empirical formula: CH_2O
CH_2O theoretical weight: 30 g/mol

NOTE: molecular weight > theoretical weight of empirical formula

$$\frac{\text{molecular weight}}{\text{empirical weight}} = \frac{180 \text{ g/mol}}{30 \text{ g/mol}} = \frac{6}{1} = 6$$

molecular formula = 6 x empirical formula

molecular formula = 6 x (CH_2O) = $C_6H_{12}O_6$

molecular formula: $C_6H_{12}O_6$ (glucose)

EXAMPLE 5:

Another unknown compound was analyzed and found to have the following percent composition: 42.9% C, 2.4% H, 16.7% N, 38.1% O. The molecular weight of the compound is known to be 168 g/mol. Find the **molecular formula** of this compound.

SOLUTION:

Remember, in order to find the molecular formula, one must first determine the **empirical formula** and then compare its theoretical weight to the known molecular weight of the compound. The **empirical formula** is determined as follows:

atom	percent	100 g sample	moles	ratio
C	42.9%	42.9 g	3.58	3
H	2.4%	2.4 g	2.40	2
N	16.7%	16.7 g	1.19	1
O	38.1%	38.1 g	2.38	2

empirical formula: $C_3H_2NO_2$

molecular weight: 168 g/mol
$C_3H_2NO_2$ theoretical weight: 84 g/mol

NOTE: molecular weight > theoretical weight of empirical formula

$$\frac{\text{molecular weight}}{\text{empirical weight}} = \frac{168 \text{ g/mol}}{84 \text{ g/mol}} = \frac{2}{1} = 2$$

molecular formula = 2 x empirical formula = 2 x ($C_3H_2NO_2$)

molecular formula: $C_6H_4N_2O_4$ (dinitrobenzene)

V. Stoichiometry

Limiting Reagent
Molarity and Stoichiometry
Enthalpy and Stoichiometry

The ability to work stoichiometry problems is one of the most important skills that a chemistry student must master. Stoichiometry is the keystone that supports the study of all chemical reactions. Thus, it is imperative that a student understand stoichiometry before he can begin to pursue further study in chemistry.

Some students become distressed just from hearing the word "stoichiometry". But these are usually the students who do not fully understand what the word means. **Stoichiometry** can be defined as a simple arithmetic approach to analyzing chemical reactions. All stoichiometry problems require the student to use a balanced chemical equation to predict the change in the amount of one substance in response to a change in the amount of another substance. For example, based on the balanced chemical equation, a student may be asked to determine the amount of *product formed* from a given amount of *reactant used.* Alternatively, the student may be asked to determine the amount of *reactant used* to produce a given amount of *product formed.* Furthermore, if the chemical equation involves more than one reactant, the student may be required to determine the amount consumed of one reactant in response to a given amount consumed of the other reactant. Similarly, if the chemical equation involves the formation of more than one product, the student may be asked to predict the amount produced of one product based on a given amount produced of the other product. Another problem may require the student to determine the change in heat associated with a chemical reaction based on the changes in the amounts of products and reactants. All of these types of problems are examples of stoichiometry. However, although stoichiometry problems can assume many different forms, the solutions to these problems are all very similar.

It is important to note that, in every case, the student **must** begin with the **balanced chemical equation.** In problems where the balanced chemical equation is not provided, the student must write the balanced equation himself before he can proceed. This tutorial assumes that the reader has had prior experience in balancing chemical equations and is aware that atoms are neither lost nor created in the process of a chemical reaction. Thus, the number of each type of atom must be equal on both the reactant and product sides of the chemical equation. Study the following examples very carefully and note the similarities between the solutions to each problem.

EXAMPLE 1:

Ammonia is produced according to the following reaction:

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

a) How many moles of ammonia would be formed when 2.00 moles of nitrogen are consumed?
b) How many moles of hydrogen are consumed when 1.20 moles of ammonia are formed?
c) How many moles of nitrogen are consumed when 0.36 mole of hydrogen is consumed?

SOLUTIONS:

The coefficients of the *balanced* chemical equation are the key to solving these types of problems. The coefficients in this chemical equation tell us the following information:

- 2 moles of ammonia are formed for every 3 moles of hydrogen consumed.
- 2 moles of ammonia are formed for every 1 mole of nitrogen consumed.
- 3 moles of hydrogen are consumed for every 2 moles of ammonia produced.
- 1 mole of nitrogen is consumed for every 2 moles of ammonia produced.
- 3 moles of hydrogen are consumed for every 1 mole of nitrogen consumed.
- 1 mole of nitrogen is consumed for every 3 moles of hydrogen consumed.

a) According to the coefficients in the balanced chemical equation, 2 moles of ammonia are formed for every 1 mole of nitrogen consumed. The following stoichiometric calculation can then be performed to determine the number of moles of ammonia formed when 2.00 moles of nitrogen are consumed:

$$2.00 \text{ moles } N_2 \text{ consumed} \times \frac{2 \ NH_3 \text{ formed}}{1 \ N_2 \text{ consumed}} = 4.00 \text{ moles } NH_3 \text{ formed}$$

b) According to the coefficients, 3 moles of hydrogen are consumed for every 2 moles of ammonia formed:

$$1.20 \text{ moles } NH_3 \text{ formed} \times \frac{3 \ H_2 \text{ consumed}}{2 \ NH_3 \text{ formed}} = 1.80 \text{ moles } H_2 \text{ consumed}$$

c) According to the coefficients, 1 mole of nitrogen is consumed for every 3 moles of hydrogen consumed (i.e. one-third as much nitrogen is used):

$$0.36 \text{ mole } H_2 \text{ consumed} \times \frac{1 \; N_2 \text{ consumed}}{3 \; H_2 \text{ consumed}} = 0.12 \text{ mole } N_2 \text{ consumed}$$

It is important to note that, in the above example, 'moles' were used as the units throughout the entire stoichiometry problem. Because the coefficients in a balanced chemical equation represent *molar ratios*, all stoichiometric calculations *must* be performed in *moles*. As a result, whenever a problem expresses given quantities in any units other than moles, the quantities must be converted to moles before the stoichiometric calculations can be performed. Furthermore, the results of stoichiometric calculations will be expressed in moles unless otherwise converted. Thus, when a problem requires an answer to be expressed in any units other than moles, the stoichiometric calculations must first be performed in moles, and the result then converted to the desired units. The most common example of this type of stoichiometry problem requires an initial conversion from mass to moles and then a final conversion from moles back to mass. A similar problem may express the given quantity in "molecules" or require the answer to be expressed in "molecules". In every case, the stoichiometric calculation that uses coefficients from the balanced chemical equation must be performed in *moles*.

EXAMPLE 2:

The combustion of glucose ($C_6H_{12}O_6$) that takes place in living cells can be expressed by the following chemical equation:

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$$

a) Calculate the number of moles of carbon dioxide that form from the combustion of 20.0 g of glucose.
b) Calculate the mass of water that would be formed from the combustion of 20.0 g of glucose.
c) Find the mass of glucose burned when 100 g of water is formed.
d) How many molecules of oxygen are required to burn 1.0 mg of glucose?

SOLUTIONS:

a)

$$20.0\ g\ C_6H_{12}O_6 \times \frac{1\ mole}{180\ g} = 0.111\ mole\ C_6H_{12}O_6 \text{ consumed}$$

$$0.111\ mole\ C_6H_{12}O_6 \times \frac{6\ CO_2}{1\ C_6H_{12}O_6} = 0.666\ mole\ CO_2 \text{ formed}$$

b)

$$20.0\ g\ C_6H_{12}O_6 \times \frac{1\ mole}{180\ g} = 0.111\ mole\ C_6H_{12}O_6 \text{ consumed}$$

$$0.111\ mole\ C_6H_{12}O_6 \times \frac{6\ H_2O}{1\ C_6H_{12}O_6} = 0.666\ mole\ H_2O \text{ formed}$$

$$0.666\ mole\ H_2O \times \frac{18\ g}{mole} = 12.0\ g\ H_2O \text{ formed}$$

c)

$$100\ g\ H_2O \times \frac{1\ mole}{18\ g} \times \frac{1\ C_6H_{12}O_6}{6\ H_2O} \times \frac{180\ g}{1\ mole} = 167\ g\ C_6H_{12}O_6 \text{ burned}$$

d)

$$1.0\ mg\ C_6H_{12}O_6 \times \frac{1\ g}{1000\ mg} \times \frac{1\ mole}{180\ g} \times \frac{6\ O_2}{1\ C_6H_{12}O_6} = 3.33 \times 10^{-5}\ mole\ O_2$$

$$3.33 \times 10^{-5}\ mole\ O_2 \times \frac{6.02 \times 10^{23}\ molecules}{1\ mole} = 2.01 \times 10^{19}\ molecules\ O_2$$

EXAMPLE 3:

Consider the following reaction:

$$Al_2(SO_4)_3 + 3Pb(NO_3)_2 \longrightarrow 2Al(NO_3)_3 + 3PbSO_4$$

a) Calculate the mass of lead sulfate formed when 5.0 moles of $Al_2(SO_4)_3$ are consumed.

SOLUTION:

$$5.0\ moles\ Al_2(SO_4)_3 \times \frac{3\ PbSO_4}{1\ Al_2(SO_4)_3} \times \frac{303\ g}{mole} = 4545\ g\ PbSO_4$$

b) What mass of lead nitrate is consumed in the question above?

SOLUTION:

$$5.0\ moles\ Al_2(SO_4)_3 \times \frac{3\ Pb(NO_3)_2}{1\ Al_2(SO_4)_3} \times \frac{331\ g}{mole} = 4965\ g\ Pb(NO_3)_2$$

c) What mass of aluminum nitrate is formed in the consumption of 100 g of $Pb(NO_3)_2$?

SOLUTION:

$$100\ g\ Pb(NO_3)_2 \times \frac{mole}{331\ g} \times \frac{2\ Al(NO_3)_3}{3\ Pb(NO_3)_2} \times \frac{213\ g}{mole} = 42.9\ g\ Al(NO_3)_3$$

d) What mass of $Pb(NO_3)_2$ is required to produce 50.0 g of $PbSO_4$?

SOLUTION:

$$50.0\ g\ PbSO_4 \times \frac{mole}{303\ g} \times \frac{3\ Pb(NO_3)_2}{3\ PbSO_4} \times \frac{331\ g}{mole} = 54.6\ g\ Pb(NO_3)_2$$

e) Determine the number of molecules of $PbSO_4$ that are formed when 5.0 moles of $Pb(NO_3)_2$ are consumed.

SOLUTION:

$$5.0\ moles\ Pb(NO_3)_2 \times \frac{3\ PbSO_4}{3\ Pb(NO_3)_2} \times \frac{6.02 \times 10^{23}}{mole} = 3.0 \times 10^{24}\ molecules\ PbSO_4$$

Limiting Reagent

The coefficients that appear before the *reactants* in a chemical equation may play an especially important role in certain stoichiometry problems. These coefficients indicate the relative molar increments by which each reactant is consumed. For example, consider the following equation:

$$A + 2B \longrightarrow C$$

The coefficients that appear before the reactants in this equation indicate that 2 moles of species B are consumed for every 1 mole of species A that is consumed. In other words, whenever 1 mole of species A reacts, 2 moles of species B must react. Thus, a *minimum* of 2 moles of species B are required to consume 1 mole of species A.

In order to further illustrate this point, let us refer back to EXAMPLE 1 which involved the following reaction:

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

In part (a) of EXAMPLE 1, we determined that 4.00 moles of NH_3 would be formed when 2.00 moles of nitrogen are consumed. Based on the coefficients of the reactants in this chemical equation, what *minimum* amount of *hydrogen* must be present for 2.00 moles of nitrogen to be consumed?

$$2.00 \text{ moles } N_2 \times \frac{3\ H_2}{1\ N_2} = 6.00 \text{ moles } H_2 \text{ required}$$

A *minimum* of 6.00 moles of hydrogen are required to consume 2.00 moles of nitrogen in the formation of the 4.00 moles of NH_3. But suppose there are more than 6.00 moles present when this reaction takes place. For example, suppose there are *7.00* moles of hydrogen present to react with the 2.00 moles of nitrogen. Would all 7.00 moles of hydrogen be consumed? Would the additional hydrogen cause a greater amount of NH_3 to be produced? The answer to both of these questions is **no**. In order to consume the additional hydrogen and to produce a greater amount of product, the number of moles of nitrogen must also be increased. As long as there are only 2.00 moles of nitrogen present, only 6.00 moles of hydrogen can be consumed. The additional mole of hydrogen would simply remain unreacted. In other words, there would be 1.00 mole of excess hydrogen when 7.00 moles of hydrogen are allowed to react with only 2.00 moles of nitrogen according to the above equation. In this case, *nitrogen* would be called the **limiting reagent**, because it is the nitrogen that is used up first. As a result, it is the amount of nitrogen present that ultimately determines the amount of NH_3 that can be produced, since no more product can be formed after the nitrogen has been used up. Thus, the **limiting reagent** is the reactant that is completely consumed in a chemical reaction and therefore determines the amount of product that will be produced.

EXAMPLE 4:

Suppose that it is your job to assemble tricycles for a toy store. When the shipment of tricycle parts arrives, you discover that the box contains the same number of wheels as tricycle frames. What will limit the number of tricycles that you can build?

SOLUTION:

Because each tricycle must have three wheels and one frame, the assembly of tricycles requires three times as many wheels as frames. If the box contains an equal number of each, you will run out of wheels before you run out of frames. Thus, the number of wheels will limit the number of tricycles that you can build.

If the assembly of tricycles in the example above is analogous to a chemical reaction, the wheels and frames represent the 'reactants' in the chemical reaction. If you are supplied with an equal number of wheels and frames, you will exhaust the supply of wheels before you will use all of the frames. Thus, the wheels would be considered the 'limiting reagent' in the process. Remember that the **limiting reagent** is completely consumed and therefore determines the amount of product that can be formed. Now let us consider a similar situation for the production of ammonia described in EXAMPLE 1.

EXAMPLE 5:

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

Suppose you wish to produce ammonia by the reaction above. You have been supplied with 30 moles of hydrogen and 30 moles of nitrogen. a) What will determine the amount of ammonia that you can produce from these reagents? b) Calculate the number of moles of ammonia that you can produce.

SOLUTIONS:

a) Because the production of ammonia requires three moles of hydrogen for each mole of nitrogen, the supply of hydrogen will be depleted before the supply of nitrogen. Thus, the moles of available hydrogen will determine the amount of ammonia that you can produce. (i.e. Hydrogen is the **limiting reagent.**)

b) limiting reagent = H_2

$$30 \text{ moles } H_2 \times \frac{2\ NH_3}{3\ H_2} = 20 \text{ moles } NH_3 \text{ produced}$$

EXAMPLE 6:

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

60 grams of hydrogen are allowed to react with 60 grams of nitrogen. a) Determine the limiting reagent. b) Calculate the mass of NH_3 that will be formed.

SOLUTIONS:

a) Because the limiting reagent depends on a comparison of *molar* ratios, the first step is to convert the given quantities to moles:

$$60\ g\ H_2 \times \frac{1\ mole}{2\ g} = 30\ moles\ H_2$$

$$60\ g\ N_2 \times \frac{1\ mole}{28\ g} = 2.1\ moles\ N_2$$

The next step is to compare the relative numbers of *available* reactants to the relative numbers of *required* reactants:

<u>AVAILABLE MOLAR RATIO</u>	vs.	<u>REQUIRED MOLAR RATIO</u>
$\frac{30\ H_2}{2.1\ N_2} = \frac{14.3\ H_2}{1\ N_2}$		$\frac{3\ H_2}{1\ N_2}$

Based on the coefficients in the chemical equation, 3 moles of hydrogen are required to react with each mole of nitrogen. According to the molar ratio of reactants *available*, 14.3 moles of hydrogen are present for every mole of nitrogen. Thus, there is more than enough hydrogen present to react with all of the available nitrogen. As a result, excess hydrogen will remain after the nitrogen is completely consumed. Nitrogen is therefore the limiting reagent and will determine the amount of product that can be formed.

b) limiting reagent: N_2

$$2.1\ moles\ N_2 \times \frac{2\ NH_3}{1\ N_2} \times \frac{17\ g}{mole} = 71.4\ g\ NH_3$$

EXAMPLE 7:

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

25 grams of hydrogen are allowed to react with 75 grams of nitrogen. a) Identify the limiting reagent. b) Calculate the mass of NH_3 that will be formed.

SOLUTIONS:

a) Remember that the first step is always to convert the given quantities to *moles*:

$$25\ g\ H_2 \times \frac{1\ mole}{2\ g} = 12.5\ moles\ H_2$$

$$75\ g\ N_2 \times \frac{1\ mole}{28\ g} = 2.7\ moles\ N_2$$

The next step is to compare the relative numbers of *available* reactants to the relative numbers of *required* reactants:

AVAILABLE MOLAR RATIO		vs.	REQUIRED MOLAR RATIO
$\frac{12.5\ H_2}{2.7\ N_2}$ =	$\frac{4.6\ H_2}{1\ N_2}$		$\frac{3\ H_2}{1\ N_2}$

The coefficients in the chemical equation indicate that 3 moles of hydrogen are required to react with every mole of nitrogen. However, according to the molar ratio of reactants *available*, 4.6 moles of hydrogen are present for every mole of nitrogen. Thus, there is more than enough hydrogen present to react with all of the available nitrogen. As a result, hydrogen will remain in excess after the nitrogen is completely consumed. Nitrogen is therefore the limiting reagent and will determine the amount of product that can be formed.

SOLUTION:

b) limiting reagent: N_2

$$2.7 \text{ moles } N_2 \times \frac{2\ NH_3}{1\ N_2} \times \frac{17 \text{ g}}{\text{mole}} = 91.8 \text{ g } NH_3$$

Molarity and Stoichiometry

Stoichiometry problems may also express given quantities in terms of concentration. Such problems generally involve reactions that take place in an aqueous solution. The most common expression of concentration is **molarity**. Molarity is defined as the number of *moles* of solute (i.e. substance dissolved) per liter of *solution*. Stoichiometry problems that employ the units of molarity are very similar to those that express the given quantities in grams. Remember that the coefficients in a balanced chemical equation represent *molar ratios*. Therefore, all stoichiometric calculations that use coefficients from the balanced equation must be performed in *moles*. Whether a problem expresses the given quantities in grams or molarity, the first step is always to convert these quantities to *moles*. The stoichiometric calculations that follow are identical. Furthermore, a problem that requires the answer to be expressed in molarity is similar to a problem that requires the answer to be expressed in grams. In either case, the stoichiometric calculations are first performed in moles, and the result then converted to the necessary units. This section is therefore little more than an extension of the *units* that may be involved in a stoichiometry problem. Only the conversion factors change. The general approach to the problem itself remains the same.

Before a student can begin to work stoichiometry problems that involve solutions, he must be able to use **molarity** to convert between liters and moles. A student who feels uncomfortable working with the units of molarity should work the following problem before proceeding to the examples that combine molarity with stoichiometry.

EXAMPLE 8: MOLARITY REVIEW

Beaker A contains 400 mL of a 0.50 M solution of sodium chloride. Beaker B contains 100 mL of a 2.00 M solution of potassium nitrate.

a) Determine the number of moles of NaCl present in Beaker A.
b) Determine the number of *grams* of KNO_3 dissolved in Beaker B.
c) What volume of the solution in Beaker A contains 1.0 g of sodium chloride?
d) Suppose the contents of Beaker B are added to Beaker A. What is the concentration of potassium nitrate in the resulting solution?

SOLUTIONS:

a)

$$400 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.50 \text{ moles}}{\text{L}} = 0.20 \text{ moles NaCl}$$

b)

$$0.100 \text{ L} \times \frac{2.00 \text{ moles}}{\text{L}} \times \frac{101 \text{ g}}{\text{mole}} = 20.2 \text{ g } KNO_3$$

c)

$$1.0 \text{ g} \times \frac{1 \text{ mole}}{58.5 \text{ g}} \times \frac{1 \text{ L}}{0.5 \text{ moles}} = 0.034 \text{ L} = 34 \text{ mL}$$

d) NOTE: The new volume is 500 mL or 0.500 L, but the number of moles of potassium nitrate has not changed.

$$0.100 \text{ L} \times \frac{2.00 \text{ moles}}{\text{L}} = 0.200 \text{ moles } KNO_3$$

$$\frac{0.200 \text{ moles}}{0.500 \text{ L}} = 0.400 \text{ M in } KNO_3$$

EXAMPLE 9:

Let us now return to the reaction given in EXAMPLE 3:

$$Al_2(SO_4)_3(aq) + 3Pb(NO_3)_2(aq) \longrightarrow 2Al(NO_3)_3(aq) + 3PbSO_4(s)$$

NOTE: Although the questions in this example will incorporate the units of **molarity**, the strategy of the stoichiometry remains the same as that used in EXAMPLE 3. Remember that this example is merely an extension of the units involved in an otherwise unchanged stoichiometry problem.

a) Calculate the mass of lead sulfate formed when **2.0 L of 2.50 M $Al_2(SO_4)_3(aq)$ are consumed.**
b) What volume of 1.20 M lead nitrate solution is consumed when **2.0 L of 2.50 M $Al_2(SO_4)_3(aq)$ are consumed.**
c) What volume of 2.50 M $Al_2(SO_4)_3(aq)$ is required to produce 50.0 g of $PbSO_4$.

SOLUTIONS:

a)

$$2.0\ L \times \frac{2.50\ moles}{L} = 5.0\ moles\ Al_2(SO_4)_3\ consumed$$

$$5.0\ moles\ Al_2(SO_4)_3 \times \frac{3\ PbSO_4}{1\ Al_2(SO_4)_3} \times \frac{303\ g}{mole} = 4545\ g\ PbSO_4$$

b)

$$5.0\ moles\ Al_2(SO_4)_3 \times \frac{3\ Pb(NO_3)_2}{1\ Al_2(SO_4)_3} = 15.0\ moles\ Pb(NO_3)_2\ consumed$$

$$15.0\ moles \times \frac{1\ L}{1.20\ moles} = 12.5\ L\ of\ 1.20\ M\ Pb(NO_3)_2\ solution\ consumed$$

c)

$$50.0\ g \times \frac{1\ mole}{303\ g} = 0.165\ mole\ of\ PbSO_4\ to\ be\ produced$$

$$0.165\ mole\ PbSO_4 \times \frac{1\ Al_2(SO_4)_3}{3\ PbSO_4} \times \frac{1\ L}{2.5\ moles} = 0.022\ L = 22\ mL\ required$$

EXAMPLE 10:

Consider the following reaction that takes place when 70 mL of a 1.0 M carbonic acid solution are mixed with 30 mL of a 6.0 M sodium hydroxide solution:

$$H_2CO_3(aq) + 2NaOH(aq) \longrightarrow 2H_2O + Na_2CO_3(aq)$$

a) Determine the number of moles of water that are formed when these two solutions are mixed.

b) What is the concentration of sodium carbonate in the resulting solution?

SOLUTIONS:

a) Remember that when the available amounts of *both* reactants are given, we must first identify the **limiting reagent** (i.e. the reactant that is consumed first). The available *moles* of this reagent will determine the moles of water formed.

$$70 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.0 \text{ mole}}{\text{L}} = 0.07 \text{ mole } H_2CO_3 \text{ available}$$

$$30 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{6.0 \text{ moles}}{\text{L}} = 0.18 \text{ mole NaOH available}$$

AVAILABLE MOLAR RATIO vs. REQUIRED MOLAR RATIO

$$\frac{0.18 \text{ NaOH}}{0.07 \text{ } H_2CO_3} = \frac{2.6 \text{ NaOH}}{1 \text{ } H_2CO_3} \qquad \frac{2 \text{ NaOH}}{1 \text{ } H_2CO_3}$$

limiting reagent: $\mathbf{H_2CO_3}$ (NaOH is present in excess)

$$0.07 \text{ mole } H_2CO_3 \times \frac{2 \text{ } H_2O}{1 \text{ } H_2CO_3} = \boxed{0.14 \text{ mole } H_2O \text{ formed}}$$

b) In order to answer this question we must calculate the total volume of the final solution and the number of *moles* of sodium carbonate formed in the reaction.

total volume = 70 mL + 30 mL = 100 mL = 0.1 L

$$0.07 \text{ mole } H_2CO_3 \times \frac{1 \text{ } Na_2CO_3}{1 \text{ } H_2CO_3} = 0.07 \text{ mole } Na_2CO_3 \text{ formed}$$

$$\frac{0.07 \text{ mole}}{0.1 \text{ L}} = \boxed{0.7 \text{ M in sodium carbonate}}$$

Enthalpy and Stoichiometry

Although we have already explored many different versions of the stoichiometry problem, there remains an additional aspect of stoichiometry that we have yet to encounter. Stoichiometry problems may also involve the principles of **thermochemistry**. Such problems incorporate the concept of **enthalpy** (H) as it relates to a chemical reaction. A characteristic change in enthalpy is usually associated with every chemical process. The **enthalpy change** (ΔH) for a chemical reaction may be thought of as the flow of heat that takes place in the course of the reaction at constant pressure. Recall that a chemical reaction may be described as endothermic or exothermic depending of the *direction* of heat flow. An **endothermic** process is accompanied by a flow of heat *into* the system, whereas an **exothermic** process is accompanied by a flow of heat *out of* the system. An endothermic reaction therefore *absorbs* heat and is characterized by a *positive* ΔH, whereas an exothermic reaction *releases* heat and is characterized by a *negative* ΔH.

The *quantity* of heat (expressed in **joules**) that is transferred in a chemical reaction depends not only on the characteristic enthalpy change of that particular reaction, but also on the amounts of reactants consumed and products formed. As the amounts of reactants and products increase, the quantity of heat transferred in the course of the reaction also increases. Furthermore, the change in the quantity of heat that is either absorbed or released is related to the coefficients of the balanced chemical equation. The quantity of heat transferred can therefore be determined from the characteristic enthalpy change, the coefficients of the balanced chemical equation, and the amounts of reactants consumed or products formed. As a result, many stoichiometry problems that incorporate thermochemistry require the student to calculate the amount of heat absorbed or released based on a given amount of reactant consumed or product formed. Similar problems may ask the student to determine the amount of reactant consumed or product formed based on the amount of heat absorbed or released.

Many students find it helpful to express the heat *absorbed* in an endothermic reaction as another "reactant" in the chemical equation. Similarly, the heat *released* in an exothermic reaction may be expressed as another "product" in the chemical equation. For example, consider the following chemical equations:

$$A + 2B \longrightarrow C \qquad \Delta H = 1000\ kJ \quad \text{(endothermic)}$$
$$C + 6D \longrightarrow 6E \qquad \Delta H = -500\ kJ \quad \text{(exothermic)}$$

These chemical equations may be written to incorporate the enthalpy change into the equation itself as described above:

$$1000\ kJ + A + 2B \longrightarrow C$$
$$C + 6D \longrightarrow 6E + 500\ kJ$$

When the equations are written to include heat as another reactant or product, the relationship between the coefficients and the quantity of heat transferred becomes more obvious.

It is important for the student to realize that stoichiometry problems that involve enthalpy changes are very similar to all other stoichiometry problems. The stoichiometric calculations involved in these problems are identical to those previously encountered. These problems are merely an extension of the information involved in the stoichiometric calculations. The general approach to solving these problems again remains the same.

EXAMPLE 11:

Recall the combustion of glucose from EXAMPLE 2:

$$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

This reaction has a characteristic ΔH of -2802 kJ at 25°.

a) Calculate the amount of heat that is evolved in the combustion of 20.0 g of glucose as described in EXAMPLE 2.
b) Calculate the heat evolved in the formation of 100.0 g of water by this reaction.

SOLUTIONS:

The negative ΔH value indicates that the combustion of glucose is an exothermic process. We should also know from experience that heat is evolved whenever something burns. We can therefore write the chemical equation to include 2802 kJ of heat as another product:

$$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O + 2802\ kJ$$

When the chemical equation is written to include the heat released as a product, we can more easily interpret the following information from the coefficients:

- **2802 kJ of heat are released for every *mole* of glucose burned.**
- **2802 kJ of heat are released for every 6 *moles* of oxygen consumed.**
- **2802 kJ of heat are released for every 6 *moles* of carbon dioxide formed.**
- **2802 kJ of heat are released for every 6 *moles* of water formed.**

a)

$$20.0\ g\ glucose \times \frac{1\ mole}{180\ g} \times \frac{2802\ kJ\ heat}{mole\ glucose} = 311\ kJ\ heat\ evolved$$

b)

$$100.0\ g\ water \times \frac{1\ mole}{18\ g} \times \frac{2802\ kJ\ heat}{6\ moles\ water} = 2594\ kJ\ heat\ evolved$$

EXAMPLE 12:

Natural gas (methane) burns according to the following reaction:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H = -891 \text{ kJ}$$

a) How much heat is produced by the combustion of 20.0 g of methane?
b) What mass of methane is required to produce 1000 joules of heat?

SOLUTIONS:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) + 891 \text{ kJ}$$

- 891 kJ of heat are released for every mole of methane burned.

a)
$$20.0 \text{ g } CH_4 \times \frac{1 \text{ mole}}{16 \text{ g}} \times \frac{891 \text{ kJ}}{\text{mole } CH_4} = 1114 \text{ kJ heat produced}$$

b)
$$1000 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{1 \text{ mole } CH_4}{891 \text{ kJ}} \times \frac{16 \text{ g}}{\text{mole}} = 0.02 \text{ g methane}$$

EXAMPLE 13:

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g) \qquad \Delta H = -92 \text{ kJ}$$

a) In EXAMPLE 6, 60 g of hydrogen were allowed to react with 60 g of nitrogen to according to the reaction above. Determine the amount of heat that was evolved in the reaction.
b) How much heat would be evolved if 10 g of hydrogen were allowed to react with 50 g of nitrogen?

SOLUTIONS:

a) We determined in **EXAMPLE 6** that **nitrogen** is the **limiting reagent** when 60 g (30 moles) of hydrogen are allowed to react with 60 g (2.1 moles) of nitrogen. Therefore, the number of moles of nitrogen available will determine the amount of heat evolved in the reaction.

$$2.1 \text{ moles } N_2 \times \frac{92 \text{ kJ}}{\text{mole } N_2} = 193 \text{ kJ heat evolved}$$

b)
$$10 \text{ g} \times \frac{1 \text{ mole}}{2 \text{ g}} = 5.0 \text{ moles } H_2 \text{ available}$$

$$50 \text{ g} \times \frac{1 \text{ mole}}{14 \text{ g}} = 3.6 \text{ moles } N_2 \text{ available}$$

AVAILABLE MOLAR RATIO vs. **REQUIRED MOLAR RATIO**

$$\frac{5.0\ H_2}{3.6\ N_2} = \frac{1.4\ H_2}{1\ N_2} \qquad \frac{3\ H_2}{1\ N_2}$$

limiting reagent: **H_2** (insufficient H_2 to react all N_2)

$$5.0 \text{ moles } H_2 \times \frac{92 \text{ kJ}}{3 \text{ moles } H_2} = 153 \text{ kJ heat evolved}$$

EXAMPLE 14:
Sodium reacts violently with water as follows:

$$Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g) \qquad \Delta H = -368 \text{ kJ}$$

A piece of sodium was dropped into a large container of water. The chemical reaction that took place produced 500 joules of heat. The volume of the resulting NaOH solution was measured to be 1.9670 L. What is the concentration of the NaOH solution?

SOLUTION:

$$500 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} \times \frac{2 \text{ moles NaOH}}{368 \text{ kJ}} = 2.72 \times 10^{-3} \text{ moles NaOH formed}$$

$$\frac{2.72 \times 10^{-3} \text{ moles}}{1.9670 \text{ L}} = 1.38 \times 10^{-3} \text{ M}$$

EXAMPLE 15: FINAL PROBLEM

Hydrochloric acid and barium hydroxide solution react as follows:

$$2HCl(aq) + Ba(OH)_2(aq) \longrightarrow BaCl_2(aq) + 2H_2O(l) \quad \Delta H = -118 \text{ kJ}$$

Calculate the amount of heat evolved when 1.50 gallons of 0.600 M HCl are mixed with 0.830 gallon of 0.500 M $Ba(OH)_2$.

SOLUTION:

$$1.50 \text{ gallons} \times \frac{3.7854 \text{ L}}{\text{gallon}} \times \frac{0.600 \text{ mole}}{\text{L}} = \mathbf{3.41 \text{ moles HCl}}$$

$$0.830 \text{ gallon} \times \frac{3.7854 \text{ L}}{\text{gallon}} \times \frac{0.500 \text{ mole}}{\text{L}} = \mathbf{1.57 \text{ moles } Ba(OH)_2}$$

<u>AVAILABLE MOLAR RATIO</u> vs. <u>REQUIRED MOLAR RATIO</u>

$$\frac{3.41}{1.57} = \frac{2.17 \text{ HCl}}{1 \; Ba(OH)_2} \qquad \text{vs.} \qquad \frac{2 \text{ HCl}}{1 \; Ba(OH)_2}$$

limiting reagent: **$Ba(OH)_2$** (HCl present in excess)

$$1.57 \text{ moles } Ba(OH)_2 \times \frac{118 \text{ kJ}}{1 \text{ mole } Ba(OH)_2} = \mathbf{402 \text{ kJ heat evolved}}$$