<u>CHEMICAL REACTIONS AND THE MOLE CONCEPT</u>

We now turn our attention to the way the mole concept permits us to determine the amounts of substances consumed or produce chemicals by reactions.

We begin with a discussion of chemical equations and you will learn how to balance them by inspections.

Next, you will learn how to apply the mole concept to chemical equations so you perform stoichiometric calculations. You will also learn how to deal with stoichiometry of chemical reactions when they occur in solutions.

Finally you will look at aspects of stoichiometry of gases. If you've begun to develop a feel for the mole concept, the calculations you encounter in this section will not be especially difficult. If you find yourself having a lot of trouble however, go back and thoroughly review the introduction to the mole concept presented there.

<u>Chemical Reactions and Chemical Equations</u>

<u>Learn how to write and balance chemical equations by inspection</u>**.** I

Chemical equations are written to provide an overview of reactions that have either taken place or that are expected to occur. Be sure you are familiar with the terms **reactants** and **products**.

For a chemical equation to be useful in chemical calculations, it must be balanced. You should keep in mind the advice that writing a balanced equation is a two-step process.

First, write the unbalanced equation with correct formulas for each of the reactants and products.

Second, balance the equation. You must balanced the equation by inspection that is juggle the coefficients numbers preceding the formulas of each kind the same on both sides of the equation. Balance all other elements before balancing oxygen and hydrogen.

Remember once you have written the **correct formulas** for the reactants and products. You do not change the subscripts in the formula just to make the equation balance. that would change the chemical nature of the substance described by the equation. Also remember that a properly balanced equation has the smallest set of whole number coefficients. If the equation is fairly complex balance it by the ion electron method (refer to the reactions course).

EXERCISE 7: Balancing the equations below by inspection.

(a) $CuSO_4 + Al \rightarrow + A1_2(SO_4)_3 + Cu$ (b) $KClO_3 \rightarrow KCl + O_2$

(c) $CO(NH_2)_2 + H_2O \rightarrow CO_2 + NH_3$ (d) $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$

(e) $N_2O_5 + H_2O \rightarrow HNO_3$ (f) $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

Calculations Based on Chemical Equations

You **must** use a **balanced chemical equation to perform calculations involving amounts of substances involved in a chemical reaction**.

A chemical equation such as

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

tells us about reacting molecules. It gives information about what happens on a **microscopic scale**.

The mole concept allows us to expand this information up to laboratory-sized quantities (i.e a **macroscopic scale**). Whatever ratios exist among atoms or molecules of reactants and products, the same ratios exist among moles of reactants and products. For example, the equation above tells us that for every two molecules of H_2 that react. one molecule Of O_2 will also react. Scaling this to laboratory-sized quantities, we can say that for every two moles of H_2 that react one mole Of O_2 will react.

In dealing with chemical equations chemists tend to do their thinking in terms of moles. To see how this is done let's use the equation above for the reaction of H_2 with O_2 as an example.

Suppose we had a 0.40-mol sample of O_2. How much hydrogen would we need to react with it completely? The coefficients of the equation tell us that twice as many moles Of H_2 must react as O_2. Therefore we will need 0.80 mol of H_2 similarly, we can conclude that this reaction will give 0.80 mol of H_2O because the equation tells us that however many moles of H_2 are consumed, the same number of moles of H_2O will be formed.

One of the principal reasons for learning to do these calculations is so that we can work with chemical equations and chemical reactions in the laboratory. However, in the lab we can't measure moles directly; instead we can only measure mass (e.g. grams). We therefore must translate back and forth between laboratory units (grams) and chemical units (moles).

Stoichiometric Relationships

Stoichiometry and the Chemical Equation

Stoichiometry is the study of the quantitative relationships between amounts of reactants and products. A chemical reaction can be represented by the general equation.

aA + bB ® cC + dD +

REACTANTS PRODUCTS

The algebraic relationships between products and reactants in the general equation above is given by

$$\frac{n_A}{a} = \frac{n_B}{b} = \frac{n_C}{c} = \frac{n_D}{d}$$

a. b. c and **d** are called stoichiometric coefficients. For example

$4NH_3$ + $5O_2$ ® $4NO$ + $6H_2O$

$$\frac{n(NH_3)}{4} = \frac{n(O_2)}{5} = \frac{n(NO)}{4} = \frac{n(H_2O)}{6}$$

Note: **Before performing any calculations that involve a chemical equation make sure that the equation (or that part of the equation under consideration) is balanced.**

Calculations involving chemical equations

In all the conversions discussed below all that is needed is:

1. A balanced chemical equation
2. (a) Converting all quantities to moles
   (b) use the stoichiometric coefficients to get the moles of required substance (c) finally converting to other units like grams or volume etc .

Mole to mole conversions

This the simplest type of calculation where your starting materials is given in moles and the quantity of product is required in moles.

Example 10: How many moles of $H_2C_2O_4$ (oxalic acid) are required to react completely with 1.50 mole of $KMnO_4$ according to the following equation

$$2KMnO_4 + 6HCl + 5H_2C_2O_4 ® 2MnCl_2 + 10CO_2 + 2KCl + 8H_2O$$

$$n(H_2C_2O_4) = \frac{5}{2} \text{ x } n(KMnO_4)$$

$$= \frac{5}{2} \text{ x } 1.50 \text{ mol}$$

$$= 3.75 \text{ mol of } H_2C_2O_4$$

2. Gram-mole: mole-gram conversions

In this type of calculation your starting material is given in grams/moles and the quantity of the product is required in mole/grams.

Example 11: How many moles of Al_2O_3 are produced from 81.0 g of Al. The equation for the reaction is

$$Fe_2O_3 + Al ® Al_2O_3 + Fe$$

MM of Al = 27.0g mol^{-1}

Use the equation to write the reacting ratios of **Al_2O_3 to Al**

Remember that moles of Al = $\frac{g(Al)}{molarmass}$

$$\frac{n(Al_2O_3)}{1} = \frac{n(Al)}{2} (from\ eqn)$$

$$= \frac{1}{2} x \frac{g(Al)}{mm}$$

$$= \frac{1}{2} x \frac{81.0g}{27.0g\ mol^{-1}}$$

= 1.50 mol

Example 12 : The decomposition of $KClO_3$ occurs as shown below:

$$2KClO_3 ® 2KCl + 3O_2$$

Calculate the number of grams of $KClO_3$ required to produce 4.50 mol of O_2.

(MM of $KClO_3$ = 122.6g mol^{-1})

Steps in the Calculation: **1** Write the reaction ratio **2.** Express the $2KClO_3$ (the unknown) in terms of grams **3.** Substitute the known quantities

$$nKClO_3 = \frac{2}{3} nO_2$$

$$gKClO_3 = \frac{2}{3} nO_2 \times MMKClO_3$$

$$= \frac{2}{3} \times 4.50 \text{ mol} \times 122.6 \text{g mol}^{-1}$$

$$= 367.8 \text{g of } KClO_3$$

3. <u>Gram to Gram Conversion</u>

In this type of calculation your starting material is given in grams and the quantity of the required substance is needed in grams.

<u>Example 13:</u> Calculate. the number of grams of $MgCl_2$ that could be obtained from 8.50 g of HCl when the latter is reacted with excess MgO

MM (in g mol^{-1}) : HCl = 36.5 $MgCl_2$ = 95.3

MgO + HCl ® $MgCl_2$ + H_2O

Solution: 1 Check to see if the equation is balanced

2. Write down the reacting mole ratio in terms of the unknown

3. Since grams is given and required expressed the equation in terms of grams.

4. Substitute all the known values and calulate grams of $MgCl_2$

$$\frac{n_{MgCl_2}}{1} = \frac{n_{HCl}}{2}$$

$$n_{MgCl_2} = \frac{1}{2} x\, n_{HCl}$$

$$\frac{g_{MgCl_2}}{MM} = \frac{1}{2} x \frac{g_{HCl}}{MM}$$

$$g_{MgCl_2} = \frac{1}{2} x \frac{8.50 g \ of \ HCl}{36.5 g \ mol^{-1}} x\, 95.3 g \ mol^{-1} of \ MgCl_2$$

$$= 11.1 \text{g of } MgCl_2$$

<u>EXERCISE 8</u>

8.1 The air pollutant nitrogen dioxide. NO_2 dissolves in rainwater to give a dilute solution of nitric acid. The equation for the reaction is

$3 NO_2 + H_2O \rightarrow 2HNO_3 + NO$

(a) How many moles of HNO_3- are formed if 1.50 mol of NO reacts?

(b) How many moles of water are needed to react with 0.600 mol of NO_2?

(c) How many moles of NO are formed if 1.60 mol of HNO_3 are formed in the reaction?

8.2 The reaction between hydrazine. N_2H_4. and hydrogen peroxide. H_2O_2. has been used to power rockets.

$$\mathbf{N_2H_4 + 2H_2O_2 \rightarrow N_2 + 4H_2O}$$

(a) How many moles of N_2H_4 are required to react with 8.00 mol of H_2O_2?

(b) How many moles of N_2 will be formed from 8.00 mol of H_2O_2?

(c) How many moles of water will be formed from 8.00 mol of H_2O_2?

(d) How many grams of water will be formed when 3.00 mol of N_2H_4 react?

(e) How many moles of N_2 will be formed when 500 g of H_2O_2 react?

(f) How many grams of H_2O_2 are required to react with 1000 g of N_2H_4?

Limiting Reactant Calculations

In this type of calculation you are required to calculate the amount of products formed when arbitrary amounts of reactants are mixed. These calculations deal with chemical reactions in which substances are simply mixed together without prior regard for maintaining the proper mole ratios between reactants. In cases like this all the reactants usually are not consumed completely; one or more of them remains in excess. The amount of product formed in these situations is controlled by the reactant that is completely used up (**the limiting reactant**), because once it is gone no more product is able to form.

There are two steps in solving this type of problem:

1. Determine which reactant is the limiting reactant;

2 Base your calculation of the amount of product formed on the amount of the limiting reactant available.

Limiting reactant problems can be recognize because you are usually given amounts of two (or more) reactants and asked to determine the amount of a

(i) product formed. or

(ii) reactant remaining after the reaction is complete.

Example 14

Zinc and oxygen combine to produce zinc oxide according to the equation

$$2Zn + O_2 \rightarrow 2ZnO$$

(a) How many grams of ZnO will be formed if 14.3 g of Zn are mixed with 3.72 g of O_2?

(b) How many grams of excess reagent remains after the reaction is complete

Molar masses (in g mol^{-1}): O_2 = 32.0 Zn = 65.4 and ZnO = 81.4

Solution: Part a

1. You must recognize a limiting reactant problem.

2 You given the amounts of both the Zn and the O_2 therefore you can assume that it is a limiting reagent problem.

3 Check to see if the equation is balanced

4 It is best to convert all quantities to moles when working with limiting reagent problems.

4 Write down the reacting mole ratio

6 Determine the limiting reagent(in moles) by carrying out the appropriate calculations

7 Detetmine amount of product(in moles) formed using moles limiting reagent as the known quantity

8. Since grams is given and required expressed the equation in terms of grams.

$$n_{(O_2)} = \frac{g_{(O_2)}}{MM_{(O_2)}} = \frac{3.72\,g}{32.0\,g\,mol^{-1}} = 0.116\ mol$$

$$n_{(Zn)} = \frac{g_{(Zn)}}{MM_{(Zn)}} = \frac{14.3\,g}{65.4\,g\,mol^{-1}} = 0.209\ mol$$

Calculate moles of ZnO produced by available amount of Zn

From the balanced equation

$$n(ZnO) = n(Zn) = 0.209\ mol$$

Calculate moles of ZnO produced by available amount of O_2

$$\frac{n_{(ZnO)}}{2} = n_{(O_2)}$$

i.e. $n_{(Zn)} = (2 \ x\ 0.116)\ mol = 0.232\ mol$

Now compare the amounts of ZnO obtained from each of the reactants.

Amount of ZnO from zinc = 0.209 mole (by calculation)

Amount of ZnO from oxygen = 0.232 mole (by calculation)

The reactant that produces the **smaller** amount of ZnO is the **limiting reagent**.

In this case the amount of ZnO obtained from Zn is less than that obtained form oxygen.

Therefore Zn is the **limiting reagent** and the amount of ZnO actually produced in this reaction is 0.209 mol.

Expressed in grams

$$g(ZnO) = n(ZnO) \times MM(ZnO)$$

$$= 0.209 \text{ mol} \times 81.4 \text{ g mol}^{-1}$$

$$= 17.8 \text{ g}$$

This means that you have more O_2 than is needed and that some O_2 will be left over. Oxygen is called the **excess reagent**

Solution part b

1. Use information determined in part (a) as appropriate
2. In order to calculate the moles of excess reagent remaining after the reaction is complete determine the moles of excess reagent (in this case oxygen) that reacts with the limiting reagent (in this case zinc)

$$n_{(O_2)} = \frac{n_{(Zn)}}{2} = \frac{0.209 \text{ mol}}{2} = 0.105 \text{ mol}$$

Therefore

$$O_2(\text{remaining}) = O_2(\text{available}) - O_2(\text{reacted})$$

$$= 0.116 \text{ mol} - 0.109 \text{ mol}$$

$$= 0.007 \text{ mol}$$

Expressed in grams

$$= g_{O2} \text{ in excess} = 0.007 \text{ mol} \times 32.0 \text{ g mol}^{-1}$$

$$0.224 \text{ g}$$

Example 15:

A 50.0g sample of $CaCO_3$ is reacted with 35.0g of H_3PO_4

$$3CaCO_3 + 2H_3PO_4 ® Ca_3(PO_4)_2 + 3CO_2 + 3H_2O$$

(a) Determine which reactant is the limiting reagent.

(b) Calculate the mass (in grams) of $Ca_3(PO_4)_2$ formed when the reaction is complete.

(c) Calculate the mass (in grams) of excess reagent remaining after the reaction is complete

Molar masses (in g mol^{-1}): $Ca_3(PO_4)_2$ = 310.0g mol^{-1} $CaCO_3$ = 100.1 H_3PO_4 = 98.03

Convert grams of reactants to moles of reactants

$$n_{H_3PO_4} = \frac{35.0\text{g}}{98.03\text{g mol}^{-1}}$$

$$= 0.357 \text{ mol}$$

$$n_{CaCO_3} = \frac{50.0\text{g}}{100.1\text{g mol}^{-1}}$$

$$= 0.500 \text{ mol}$$

Find the limiting reagent - carry out two calculations.

(i) Calculate the amount of $Ca_3(PO_4)_2$ formed from the available amount of H_3PO_4

$$n_{Ca_3(PO_4)_2} = \frac{n_{H_3PO_4}}{2}$$

$$= \frac{0.357 \text{ mol}}{2}$$

$$= 0.179\text{mol}$$

(ii) Calculate the amount of $Ca_3(PO_4)_2$ formed from the available amount of $CaCO_3$.

$$n_{Ca_3(PO_4)_2} = \frac{n_{CaCO_3}}{3}$$

$$\frac{0.500\ mol}{3}$$

$$= 0.167 mol$$

Compare mole of $Ca_3(PO_4)_2$ obtained from moles of H_3PO_4 with that obtained from moles of $CaCO_3$. Since moles of $Ca_3(PO_4)_2$ from $CaCO_3$ is less than that from H_3PO_4, $CaCO_3$ is the <u>limiting reagent</u> and H_3PO_4 is the <u>excess reagent</u>.

(b) $\therefore$The amount of $Ca_3(PO_4)_2$ formed = 0.167mol.

g ($Ca_3(PO_4)_2$) = 0.167 mol x 310.0g mol^{-1} = 51.8 g

(c) **Determine the amount of H_3PO_4 that reacts with $CaCO_3$ (the limiting reagent).**

$$\frac{n_{H_3PO_4}}{2} = \frac{n_{CaCO_3}}{3}$$

$$n_{H_3PO_4} = \frac{2}{3}\, x\ 0.500\ mol$$

= 0.333 mol

Calculate the amount of excess reagent remaining

= n(H_3PO_4) (initially) – $n_{(H3PO4)}$(reacted with $CaCO_3$)

= 0.357 mol – 0.333 mol

= 0.0240 mol of H_3PO_4 leftover

gH_3PO_4 = nH_3PO_4 x MMH_3PO_4

= 0.0240 mol x 98.03 g mol^{-1}

= 2.35g

EXERCISE 9

For the equation.

$$N_2H_4 + 2H_2O_2 \rightarrow 4\ N_2 + 4H_2O$$

(a) How many moles of H_2O will be formed if 2.50 mol N_2H_4 are allowed to react with 4.80 mol .H_2O_2?

(b) How many moles of N_2 Will be formed if 600 g of N_2H_4 are mixed with 1200 9 Of H_2O_2?

(c) How many grams of H_2O will be produced if 83.5 g of N_2H_4 are mixed with 180 g of H_2O_2?

(d) In the mixture described in part (c). ow many grams of which reactant will be left over after the reaction has stopped?

Theoretical Yield and Percentage Yield

You should be aware that not all reactions produce the theoretical maximum amount of product. You should Know the definitions of theoretical yield and percentage yield.

The theoretical yield is the amount of product that would be produced if the reactants were to combine to the maximum extent possible. We calculate the theoretical yield from the limiting reactant by following the procedure described in the last section. In simpler cases we calculate it as the maximum amount of product formed from a particular reactant. The percentage yield is calculated as shown in the next section.

Theoretical Yield (TY). Actual Yield (AY) and Percentage Yield (PY)

TY – the maximum amount of product that can be produced in a chemical reaction.

AY – the amount of product actually obtained in the laboratory from a reaction.

e.g. consider the reaction between 2 mol of O_2 and 10 mol of H_2 to product H_2O according to the balanced equation below

	O_2	+	$2H_2$	®	$2H_2O$
I	2 mol		10 mol 0		
C	-2 mol		-4 mol		+4 mol
E	0 mol		6 mol		+4 mol

4mol of H_2O is the theoretical yield – i.e. the maximium amount of water that can be produced .

In an actual experiment only 2.5 mol of H_2O was obtained (actual yield) because of a number of factors like side reactions etc.

$$PY = \frac{AY}{TY} \times 100 \qquad \text{i.e.} \qquad PY = \frac{2.5\text{mol}}{4.0\text{mol}} \times 100$$

$$= 62.5\%$$

Note : Percent yield can be calculated using mol or grams. For the example above

2.5mol of H_2O = 45.05g of H_2O (actual yield)

4.0 mol of H_2O = 72.0g of H_2O (theoretical yield)

$$PY = \frac{AY}{TY} \times 100 \qquad = \frac{45.05g}{72.08g} \qquad = 62.5\ \%$$

Example 16 :

The gas diborane B_2H_6 can be prepared by the following reaction

$$\mathbf{3NaBH_4 + 4BF_3 \text{ ® } 3NaBF_4 + 2B_2H_6}$$

if 18.9g of $NaBH_4$ and an excess of FB_3 is reacted and 7.50g of B_2H_6 is isolated. What is the percent yield of B_2H_6? Molar masses (in g mol^{-1}): $NaBH_4$ = 37.8 B_2H_6 = 27.7

Calculate the TY

(Note if TY = AY then PY = 100%)

$$\frac{n_{B_2H_6}}{2} = \frac{n_{NaBH_4}}{3}$$

$$n_{B2H6} = \frac{2}{3} \times n_{NaBH4}$$

$$g_{B2H6} = \frac{2}{3} \times n_{NaBH4} \times MM_{B2H6}$$

$$= \frac{2}{3} \times \frac{18.9\text{g}}{37.8\text{g mol}} \times 27.7\text{g mol}^{-1}$$

= 9.23g = TY

AY = 7.50g (mass actually obtained in the laboratory).

$$Py = \frac{AY}{TY} \times 100 = \frac{7.50g}{9.23g} x\ 100$$

= 81.3%

EXERCISE 10

Glucose. $C_6H_{11}O_6$. is converted to ethyl alcohol. C_2H_5OH. and CO_2 by fermentation.

$$C_6H_{11}O_6 \rightarrow 2\ C_2H_5OH + 2CO_2$$

Starting with 200 g of glucose:

(a) What is the theoretical yield of ethyl alcohol?

(b) If 97.3 9 of C_2H_5OH was obtained what was the percentage yield?

Limiting reactant is the reactant that is completely consumed in a chemical reaction. It is the reactant that limits the amount of products that can be formed in a particular experiment.

Excess reagent is the reagent that is left over after the reaction is complete.

Side reactions: Reactions that occur in a reaction mixture other than the desired one.

Side products. Products formed in a reaction mixture other than the desired products.

Actual yield. The actual amount of a product obtained in a particular chemical reaction when the experiment is performed in the laboratory.

Theoretical yield. The maximum amount of a product that could be formed from a particular mixture of reactants.

Percentage yield. Amount of required product obtained expressed as a percentage.

EXERCISE 11 : **Limiting Reagent and percentage yield**

11.1 Aluminium chloride is made by reacting scrap aluminium with chlorine gas.

a) Write a balanced equation for the reaction.

b) Which reactant is limiting if 2.70g of aluminium and 4.05g of chlorine gas are mixed?

c) What mass of aluminium chloride can be produced? **5.08g**

d) What mass of excess reactant remains when the reaction is complete?

1.67g

11.2 Phosphorous pentachloride reacts with water to give phosphoric acid and hydrogen chloride according to the following equation.

$$\mathbf{PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl}$$

In one experiment 0.360 mol of PCl_5 was slowly added to 2.88 mol of water.

(a) Which reactant was the limiting reagent?

(b) How much of excess reagent remains of the reaction is complete?

(c) Calculate the theoretical yield of H_3PO_4.

(d) What is the percentage yield of H_3PO_4 if 1.44 mol of the acid is obtained?

11.3 The reaction of 13.0 g C_4H_9OH, 21.6 g NaBr and 33.8 H_2SO_4 yields 16.8 g C_4H_9Br in the reaction

$$\mathbf{C_4H_9OH + NaBr + H_2SO_4 \rightarrow C_4H_9Br + NaHSO_4 + H_2O}$$

What are the:

(a) theoretical yield,

(b) actual yield,

(c) the percentage yield
of this reaction

<u>PAST EXAMINATION QUESTIONS</u>

QUESTION 1 (9 MARKS)

Phosphorous pentachloride reacts with water to give phosphoric acid and hydrochloric acid according to the equation

$$PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl$$

In an experiment 0.360 mol of PCl_5 was slowly added to 2.88 mol of water.

Molar masses (g mol^{-1}): O = 16.00, H=1.01, PCl_5 = 206, H_3PO_4 = 98

(a) Determine the limiting reagent. (Show all calculations) **(2)**

(b) Calculate the mass of excess reagent that remains after the reaction is complete. **(3)**

(c) Calculate the theoretical yield of phosphoric acid **(2)**

(d) What is the percentage yield of phosphoric acid if 1.44mol of acid is obtained? **(2)**

QUESTION 2 (8 MARKS)

Consider the reaction

$$\mathbf{MnO_2 + 4HCl \text{ ® } MnCl_2 + Cl_2 + 2H_2O}$$

A mixture of 0.860 mole of MnO_2 and 48.2 g of HCl react:

(a) Which reagent will be used up first? (Show relevant calculations).

(b) Calculate the mass of Cl_2 that will be produced

(c) Calculate the percentage yield based on the limiting reagent

Molar masses: Cl = 35.5 H = 1.01

QUESTION 3 (8 MARKS)

Methanol is produced by the reaction of carbon monoxide with hydrogen.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

A reaction is carried out in which 356g of CO is mixed with 65.0g of H_2,.

Molar masses(in g mol^{-1}): C = 12.0, O = 16.0, H = 1.01

(a) Determine the limiting reagent? (Show all calculations)

(b) What mass of excess reagent remains after the reaction is complete

REACTIONS IN SOLUTION

You must learn the meanings of some terms used in discussing solutions. You should learn the meaning of molar concentration and how to use it as a conversion factor relating amounts of solute and volumes of solutions.

Reactions are carried out in solution so the reactant particles can mingle freely and the reaction can take place rapidly. In general, in a solution the solvent is the substance whose physical state doesn't change. If two liquids are mixed, the one present in the larger amount is the solvent. When one of the components of a solution is water then H_2O is taken to be the solvent. All other components of the solution are solutes. Learn the definitions of concentrated and dilute.

The proportion of solute to solvent is specified by giving the solution's concentration. Molar concentration or molarity is a convenient concentration unit for measuring small amounts of solute that are dissolved in a solution. The units of molarity is mol L^{-1} or it is also represented by **M**:

$$molarity = \frac{moles\ of\ solute}{litre\ of\ solution}$$

To calculate molarity you need to know the number of moles of the solute and the total volume of the solution in which it is dissolved.. If you know the molarity of a solution, you can use it to calculate the number of moles of solute in a specified volume of the solution, or to calculate the volume of solution needed to contain a specified number of moles of solute. Think of molarity as a conversion factor that relates moles of solute to volume of solution. You should be able to translate a label such as 2.50 M H_2SO_4 into these factors:

2.50 mol of H_2SO_4 in 1.00 L solution

1.00 L solution contains 2.50 mol H_2SO_4

They can also be written with the volume units in milliliters.

2.50 mol H_2SO_4 in 1000 mL solution

1000 mL solution contains 2.50 mol H_2SO_4

In preparing a solution of a desired molarity, remember that the solvent (water. for example) is added to the solute until the desired final volume of the solution is reached. To prepare 500 mL of solution, for example, we do not just add 500 mL of water to the solute. First the solute is dissolved in a small amount of water (usually in a volumetric flask) and then more water is added until the final volume is 500 mL

NOTE: Terminology

Solution : a homogeneous mixture of two or more substances

A solution is made up of a solute (usually the substance in smaller amount) and solvent (usually the susstance in larger amount)

Concentration refers to number of solute particles per given amount of solution (usually volume or mass)

Methods for expressing concentration of solutions

There are many ways of expressing concentration of solutions e.g.

1.	molarity	2.	mass percent
3.	molality	4.	normality
5.	mole fraction	6.	ppm (parts per million)

Only molarity and mass percent will be considered in this course.

Example 17: A student prepared a solution of NaCl by dissolving 1.461g of NaCl in a 250 mL volumetric flask. What is the molarity of the solution?

MM : NaCl = 58.44 g mol^{-1}.

From the definition of molarity

$$M = \frac{n}{V}$$

Expressing the mole in terms of grams and molar mass(MM) and the volume in terms of litres the equation becomes

$$= \frac{g_{NaCl}}{MM} x \frac{1}{V}$$

$$= \frac{1.461\text{g}}{58.44\text{g mol}^{-1}} \text{ x } \frac{1}{0.250\text{L}}$$

= 0.1000 mol L^{-1} NaCl solution

<u>Example 18</u>: How many mL of 0.250 mol L^{-1} NaCl solution must be measured to obtain 0.100 mol of NaCl

$$M = \frac{n}{V}NaCl$$

$$V = \frac{nNaCl}{M}$$

$$= \frac{0.100\ mol}{0.250\ mol\ L^{-1}}$$

$= 0.400L$

convert to mL - x 1000

$= 0.400L \times 1000\ mL\ .L^{-1}$

= 400mL of NaCl soln. Is required.

<u>EXERCISE 12</u>

12.1 Calculate the molarities of the following solutions:

(a) 0.350 mol $NaHCO_3$ in 0.400 L of solution

(b) 0.250 mol KCl in 200 mL of solution

(c) 15.6 9 Of $MgCl_2$ in 300 mL of solution

(d) 1.85 g of $AgNO_3$ in 75.0 mL of solution

12.2 Calculate the number of moles of $CaCl_2$ in:

(a) 1.15 L of 0.840 M $CaCl_2$ solution. (b) 325 mL of 0 150 M $CaC1_2$ solution

12.3 What volume (in mL) of 3.00 M NH_3 solution contains

(a) 1.35 mol of NH_3? (b) 21.4 g of NH_3?

12.4 How many grams of KNO_3 are needed to prepare 750 mL of 0.200 M KNO_3 solution?

EXERCISE 13

13.1 Calculate the molarity of the solute in each of the following solutions

(a) 1.14 mol KI in 1.50 L of solution

(b) 0.240 mol $CaC1_2$ in 500 mL of solution

(c) 3.50 g of NaCl in 0.0500 L of solution

(d) 4.25 9 $MgSO_4$ in 75.0 mL of solution

13.2 How many moles of $KClO_3$ are in 500 mL of 0.150 M solution?

13.3 How many moles of urea are in 250 mL of urine if the urea concentration is 0.320 M?

13.4 A normal adult excretes about 1500 mL of urine per day. If the urea concentration is 0.320 M, how many grams of urea are excreted per day? Urea has the formula, $CO(NH_2)_2$

13.5 What is the molar concentration of each ion in the following solutions?

(a) 0.300 M $AlCl_3$ (b) 0. 150 M $(NH_4)_2Cr_2O_7$ (c) 0.200 M Na_3PO_4

(d) 0.400 M $Cr_2(SO_4)_3$ (e) 0 100 M $Ba(OH)_2$

13.6 Calculate the number of mL of 0.250 M $CaC1_2$ required to react completely with 300 mL of 0.150 M $AgNO_3$ according to the equation,

$$\mathbf{Ag^+ + Cl^- -+ AgCl(s)}$$

13.7 Calculate the number of moles of solid AgCl that forms if 300 mL of 0.240 M $AgNO_3$ is added to 200 mL of 0.480 M HCl. The reaction is:

$$\mathbf{Ag^+ + Cl^- \rightarrow AgCl(s)}$$

13.8 In Question 13.7 what will be the molar concentrations of the ions remaining in solution after the reaction is complete?

13.9 In an experiment a student mixed 300 mL of 0.200 M $Ba(OH)_2$ with 500 mL of 0.100 M $CuSO_4$

(a) Write net ionic equations for any chemical reactions that occur in this mixture.

(b) For any solids formed, calculate their amounts in moles.

(c) Calculate the molar concentrations of any ions that remain in solution after reaction is complete.

13.10 A 6.73g sample of Na_2CO_3 is dissolved in enough water to make 250 mL of solution.

(a) What is the molarity of the sodium carbonate?

(b) What are the molar concentrations of Na^+ and CO_3^{2-} ions?

13.11 What is the mass in grams of solute in 250 mL a 0.0125 mol L^{-1} solution of $KMnO_4$?

13.12 What volume of 0.123mol L^{-1} NaOH, in mL, contains 25g of NaOH?

Remember the following terms

Precipitate: A solid that is formed in a solution usually as the result of a chemical reaction.

Solute. A substance dissolved in a solvent.

Solvent. Generally the substance in a solution that is present in largest amount. If one substance is a liquid it is normally considered to be the solvent.

Concentrated: A large proportion of solute to solvent in a solution.

Dilute: Very little solute dissolved in a solution-a low ratio of solute to solvent.

Concentration: A quantitative statement of the proportion of solute to solvent or of solute to the total amount of solution.

Molar concentration (molarity): A ratio of moles of solute to liters of solution. It is the number of moles of solute per liter of solution.

Molar: A term that describes the molar concentration of a solute in a solution. It means moles of solute per liter of solution.

Preparing Solutions by Dilution

You must know how to perform calculations that are necessary when dilute solutions are to be prepared from concentrated solutions.

When ever H_2O is added to a solution it becomes diluted i.e. the molarity (or mass percent) decreases. The amount of solute however remains in a solution remains constant as the solution is diluted, we can use the simple relationship

Moles of solute before dilution = Moles of solute after dilution

$$n_{before} = n_{after}$$

From the definition of molarity i.e. $M = \frac{n}{V}$

Therefore moles = M x V(L)

And $M_jV_j = M_fV_f$ = n which is constant

i and **f** refer to the initial and final solution.

It is important to note that mixing two solutions containing different ions will also involve dilution

Example 1. Mixing 1L of a 1mol L^{-1} NaCl with 1L of a 1 mol L^{-1} KBr will result in the concentration of all the ions decreasing

$$V_T = 1L + 1L = 2L$$

e.g. $$M_{NaCl} = \frac{n}{V} = \frac{1mol}{2L}$$

$$= 0.5mol\ L^{-1}$$

Notice that in this case the concentration of NaCl decreased by half

Examples 19 (a)

How many mL of a 10.0 mol L^{-1} HCl solution is required to prepare 25.0L of 0.500 mol L^{-1} HCl solution?

n_{HCl} (Before) = n_{HCl} (After)

$M_{HCl}(B) \times V_{HCl}(B) = M_{HCl}(A) \times V_{HCl}(A)$

$$V_{HCl}(B) = \frac{0.500mol\ L^{-1}\ x\ 25.0L}{10.0mol\ L-1}$$

$$= 1,25L$$

$$= 1250mL$$

Example 19 (b)

If 82.5 mL of 6.25 mol L^{-1} solution is diluted to a final volume of 250 mL, What is the molarity of the dilute HCl solution?

n_{HCl} (Before) = n_{HCl} (After)

$M_{HCl}(Before) \times V_{HCl}(Before) = M_{HCl}(After) \times V_{HCl}(After)$

$$M_{HCl}(After) = \frac{M_{HCl}(Before)\ x\ V_{HCl}(Before)}{V_{HCl}(after)}$$

$$= \frac{6.25mol\,L^{-1}\ x\,0.0825L}{0.250L}$$

$$= 2.06\ mol\ L^{-1}$$

You must know how to prepare a dilute solution from a concentrated solution. Finally, for safety reasons when diluting concentrated laboratory reagents with water. always add the concentrated reagent to the water.

EXERCISE 14

14.1. How many mL of 3.00 M HCI must be used to prepare 500 mL of 0.100 M HCI?

14.2. How much water must be added to 50.0 mL of 1.00 M NaOH to produce 0.100 M NaOH?

14.3 If 4.00 mL of 0.0250 mol L^{-1} $CuSO_4$ is diluted to 10.00 mL with pure water what is the molarity of the copper salt in the solution?

14.4 If 300 mL of H_2O is added to 600 mL of 0.960 M H_2SO_4. what is the final molarity of the H_2SO_4?

Note the following:

Reagent. A term often used to refer to common chemicals that stocked in the laboratory.

Dilution. The act of making a solution less concentrated by the addition of more solvent.

Volumetric flask. A special flask designed to hold a specified volume of solution when filled to the line etched around the neck of the flask.

ANALYSIS

The Stoichiometry of Reactions in Solution using molarity as the concentration unit

You must know how to use molarity in dealing with the stoichiometry of reactions that take place in solutions.

In most ways reactions in solution are no different than those that occur elsewhere.

The stoichiometric relationships between reactants and products are determined by the mole ratios specified by the coefficients in the balanced chemical equation.

Therefore the first step in working problems involving reactions in solute is writing a properly balanced equation.

The principal way "solution stoichiometry" problems differ from others that you've done so far is that amounts of reactants and products can be specified as volumes of solutions of known concentrations. It is useful to remember that the product of molarity and volume gives moles: it is just necessary to be careful about the volume units.

$$molarity = \frac{moles\ of\ solute}{litre\ of\ solution}$$

Moles =volume x molarity

In terms of units $\quad$ Mol = L X mol L^{-1}

Mol(unit)

Example 20 : Copper is dissolved in dilute HNO_3 solution by the following reaction

$3Cu + 8HNO_3$ ® $Cu(NO_3)_2 + 2NO + 4H_2O$

How many mL of 3,00 mol L^{-1} HNO_3 solution can react with 10 g of Cu?

Molar mass of Cu = 635 g mol^{-1}

Calculate moles of Cu = $\frac{10\text{g}}{63.5\text{ g mol}^{-1}}$ = 0.157 mol Cu

Calculate moles of HNO_3

From the equation $\frac{n_{HNO3}}{8} = \frac{n_{Cu}}{3}$ $= n_{Cu} \text{ x } \frac{8}{3}$

$= \frac{8}{3} \text{ x } 0.157$ mol of HNO_3

= 0.420 mol of HNO_3

Volume of HNO_3 is calculated from

$$M = \frac{n}{V}$$

therefore $V = \frac{n}{M}$

$$= \frac{0.420\text{mol}}{3.00\text{mol L}^{-1}}$$

Example 21

How many milliliters of 0.200 M $Ba(NO_3)_2$ are needed to react completely with 75.0 mL of 0.150 M $Fe_2(SO_4)_3$ solution? The net ionic equation for the reaction is

$$\mathbf{Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4(s)}$$

First determine how many moles of sulfate ion there are in the $Fe_2(SO_4)_3$ solution. Then we can calculate how many moles of barium ion are needed, and finally we can calculate the volume of the $Ba(NO_3)_2$ solution needed.

The molar concentration of SO_4^{2-} in the $Fe_2(SO_4)_3$ solution is SO_4^{2-} concentration = 3 x (0. 150 M) = 0.450 M

Therefore, in 75.0 mL of this solution there is 0.0750 L x 0.450 mol L^{-1} = 0.0338 mol SO_4^{2-}

From the stoichiometry of the net ionic equation, it is obvious that the number of mole of Ba^{2+} needed for the reaction is also 0.0338 mol.

Moles of Ba^{2+} = Moles of SO_4^{2-} = 0.0338 mol

The volume of the $Ba(NO_3)_2$ solution required is obtained from the molarity of the solution. In this solution the Ba^{2+} concentration is 0.200 M because each formula unit of $Ba(NO_3)_2$ contains one Ba^{2+}. Therefore,

$$V_{\left(Ba^{2+}\right)} = \frac{0.0338\ mol}{0.200\ mol.L^{-1}}$$

$$V_{\left(Ba^{2+}\right)} = \frac{n}{M}$$

= 0.169 L

Convert litres to milliliters by multiplying by a thousand

= 0.169 L x 1000 mL L^{-1} = 169 mL

The volume of the barium nitrate solution needed is 169 mL.

You should remember that limiting reactant calculations applies to solution chemistry as well.

EXERCISE 15

15.1 Consider the following reaction that takes place in solution

$$\mathbf{CaCl_2(aq) + 2AgNO_3(aq) \rightarrow Ca(N0_3)_2(aq) + 2AgCl(s)}$$

Suppose that we began with 200 mL of a 0.200 M solution of $CaCl_2$

(a) How many moles of $CaCl_2$ are there in this solution?

(b) How many moles of $AgNO_3$ would be required to react completely with this amount of $CaCl_2$?

(c) How many milliliters of 0.500 M $AgNO_3$ solution would be needed to contain this number of moles of $AgNO_3$?

15.2 Consider the reaction described in Question 13 above. Suppose that 200 mL of 0.150 M $CaCl_2$ solution are mixed with 180 mL of 0.220 M $AgNO_3$ solution.

(a) How many moles of $CaCl_2$ are in the first solution?

(b) How many moles of $AgNO_3$ are in the second solution?

(c) When the solutions are mixed, which is the limiting reactant?

(d) How many moles of AgCl are formed in the reaction?

(e) How many moles of the reactant in excess are left over after the reaction has stopped?

(f) What is the molar concentration of the excess reactant in the reaction mixture after the reaction has stopped?

15.3 How many grams of Na_2CO_3 are required for complete reaction with 25.0 mL of 0.155 mol L^{-1} HNO_3. The equation for the reaction is

$$\mathbf{Na_3CO_3 + HNO_3 \rightarrow NaNO_3 + CO_2 + H_2O}$$

15.4. What mass of Na must react with 125 mL of water to produce a solution that is 0.250 mol L-1 NaOH? (Hint : write a balanced equation for the reaction between sodium and water).

15.5 A 0.2048 – gram sample of oxalic acid, $H_2C_2O_4$, requires 24,87 mL of a particular NaOH(aq) to complete the following reaction. What is the molarity of the NaOH(aq)? The equation for the reaction is

$$\mathbf{H_2C_2O_4(s) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O}$$

15.6 A 1.000-gram sample of an iron ore containing Fe_2O_3 was dissolved in acid and all the iron converted to Fe^{2+}. The solution was titrated with 90,4 mL of 0.02000 mol L^{-1} $KMnO_4$. The reaction that occurs is as follows:

$$\mathbf{5Fe^{2+} + MnO_4^{-}\ 8H^{+} \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O}$$

15.7 A 25,00 mL sample of HCl was added to a sample of $CaCO_3$. All the $CaCO_3$ reacted, leaving some some HCl(aq) over.

$$CaCO_3(s) + 2HCl \rightarrow CaCl_2(aq) + H_2O + CO_2(g)$$

The excess HCl (aq) required 45,76 mL of 0,01221 mol L^{-1} $CaCO_3(s)$ to complete the following reaction.

$$2\,HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O$$

What is the mass, in grams, of the original $CaCO_3$ sample?

CHEMICAL ANALYSIS

Your study of stoichiometry so far should have convinced you that

(a) if you know the balanced equation for a reaction occurring between two reactants, (b) if the reaction is rapid and complete,

(c) and if you know the exact quantity of one of the reactants, then you can always obtain the exact amount of any other of the substances in the reaction.

This is the essence of any technique of **quantitative chemical analysis,** the determination of the amount of a given constituent in a mixture.

Suppose, for example, you want to analyze a type of common clover for the quantity of oxalic acid, $H_2C_2O_4$, in the leaves. We know this acid reacts with the base sodium hydroxide in aqueous solution according to the balanced equation

$$H_2C_2O_4,(aq) + 2\,NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\,H_2O(l)$$

You can tell exactly how much oxalic acid is present in a given mass of clover leaves if the following conditions are met:

1. the reaction is done in such a way that you know when the sodium hydroxide being added is exactly the amount required to react with all the oxalic acid present in solution;
2. the exact volume of the base is known when you have reached the point where the exact stoichiometric reaction has occurred; and
3. the concentration of the sodium hydroxide is known exactly.

These conditions are fulfilled in **a titration,** The solution containing oxalic acid is placed in a flask along with a highly coloured dye, the purpose of which is explained below. Sodium hydroxide of exactly known concentration is placed in a **burette**, a type of measuring cylinder most commonly of 50.0 mL volume and calibrated in 0. 1 mL divisions.

As the sodium hydroxide is added slowly to the acid solution in the flask, the acid is consumed by reaction with the base according to the net ionic equation

$$\mathbf{H_3O^+(aq) + OH^-(aq) \quad 2\,H_2O(l)}$$

As long as $\mathbf{H_3O^+}$ ' from the acid is present in solution, every mole of OH^- supplied by the base from the burette is consumed by the $\mathbf{H_3O^+}$. The point at which the number of moles of OH^- added is equal to the number of moles of $\mathbf{H_3O^+}$ supplied by the acid is called the **equivalence point.** To indicate to us when this point has been reached, we added to the solution an **acid-base indicator,** the dye mentioned above. The dye is selected for its sensitivity to the $\mathbf{H_3O^+}$ concentration in solution; it undergoes a strong colour change at a hydronium ion concentration as near the equivalence point as possible.

When the equivalence point has been estimated in a titration, the volume of base used from the beginning of the titration can be determined by reading the calibrated burette. If you know the concentration of the base in units of moles/litre, you can then calculate the exact number of moles of base used from

Moles of base used = molarity of base (mol/L) x volume of base (L)

Since the balanced equation for the acid-base reaction gives you the stoichiometric factor, you can use this mole-ratio conversion factor to convert moles of base added to the exact number of moles of acid present in the original sample.

ACID-BASE TITRATIONS

Practice Exercise: Suppose you have 1.034 g of clover leaves and you extract the oxalic acid from them into a small amount of water. This solution of oxalic acid is found to require 34.47 mL of 0.100 M NaOH for titration to the equivalence point. What is the percentage by mass of oxalic acid in the leaves?

Solution: Step 1. Write the balanced equation.

$1H_2C_2O_4$, (aq) + 2 NaOH(aq) $\rightarrow$ $Na_2C_2O_4$(aq) + 2 H_2O(l

Step 2. Write an expression to show the reacting ration of acid to base in order to calculate the amount of acid required to react with the given amount of base

$$\frac{n_{acid}}{1} = \frac{n_{base}}{2}$$

$$= \frac{0.100\ mol\ L^{-1}\ x\ 0.03447\ L}{2}$$ = 0.00172 mol of oxalic acid

Step 4. Calculate the number of grams of acid in solution.

$G_{acid} = n_{acid}$ x MM_{acid}

= 0.00172 mol x 90.04 g mol^{-1}

= 0.155 g

Step 5. Calculate the weight percentage of $H_2C_2O_4$, in 1.034 g of leaves.

Weight percentage = $\frac{0.155\ \text{g of acid}}{1.034\ \text{g of leaves}}$ x 100

= 15 %

In the preceding example, the exact concentration of the base, NaOH, was known. The procedure by which the exact concentration of any reagent is determined is called **standardization**, and there are two general approaches to this as illustrated by the next example

GRAVIMETRIC CHEMICAL ANALYSIS OF MIXTURES BY PRECIPITATION REACTIONS

The second method of anaylsis involves the analysis of a substance by forming and weighing a compound of known composition that contains all of one component of the original sample being analyzed

If a mixture contains an acid, the quantity of acid present can be determined by an acid-base titration. On the other hand, if a mixture contains a substance that forms an insoluble salt, its amount can be determined by precipitating the salt quantitatively (that is, knowing that essentially all of the salt is precipitated, and a negligible amount remains in solution). As an example, suppose you wish to determine the amount of barium in a sample that contains an ionic compound of Ba^{2+} plus some impurity. A procedure for doing this is involves a weighed sample containing an unknown amount barium dissolved in water. Sulphuric acid, H_2SO_4, is added to the solution and the Ba^{2+} ion in the water and the SO_4^{2-} ion from the added acid combine to produce insoluble $BaSO_4$. If sufficient acid is added, all the Ba^{2+} that was in the unknown sample is precipitated and can be isolated by filtration; after drying, the mass of the $BaSO_4$ can be determined. Since you know that one mole of $BaSO_4$ contains one mole of Ba^{2+}, the number of moles (and grams) of Ba^{2+} in the unknown can be calculated. This type of analysis is sometimes called gravimetric analysis, which in contrast with titrations is called volumetric analysis.

Analysis of substances by precipitation reactions is an important technique, and two such analyses are explored in the following example and exercise.

Example 22(a) Suppose you have some solid that consists of some *$BaCl_2$ contaminated* with some NaCl. To analyse the mixture you must separate the Ba^{2+} from the Na^+ and then isolate the Ba^{2+} ions in the form of a compound of known formula. Therefore you take 1.023 g of the solid mixture, dissolve it in water and add H_2SO_4 to form insoluble $BaSO_4$ and leave NaCl in solution. If the $BaSO_4$ has a mass of 0.560 g after isolating and drying, calculate (a) the mass percent of barium in the sample and (b) the number of grams of $BaCl_2$ in the original mixture.

$$\mathbf{Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4(s)}$$

(a) Step 1: Determine the number of moles of of $BaSO_4$

$$n(BaSO_4) = \frac{g(BaSO_4)}{MM(BaSO_4)} = \frac{0.560\ g}{233.4\ g\ mol^{-1}} = 0.002399\ mol$$

Step 2: Determine the number of moles of of Ba

The number of moles of of Ba = The number of moles of of $BaSO_4$

= 0.002399 mol

Step 3 Determine the mass of Ba in the sample

$$g(Ba^{2+}) = n(Ba^{2+}) \text{ x } MM(Ba^{2+})$$

$$= 0.002399 \text{ mol x } 137.3 \text{ g mol}^{-1}$$

$$= 0.3294 \text{ g}$$

Step 4 Determine the mass percent of barium in the sample

$$\% \text{ Ba in the sample} = \frac{g_{Ba}}{g_{sample}} \text{ x } 100$$

$$\% \text{ Ba} = \frac{0.3294 \text{ g}}{1.023 \text{ g}} \text{ x } 100$$

$$= 32.16 \%$$

(b) Step 5 Determine the number of moles of of $BaCl_2$

no of moles of $BaCl_2$ = no of moles of $BaSO_4$

= 0.002399 mol

step 6 Determine the mass of $BaCl_2$

$$g(BaCl_2) = n(BaCl_2) \text{ X } MM(BaCl_2)$$

$$= 0.002399 \text{ mol x } 208.2 \text{ g mol}^{-1}$$

= 0.4995 g of $BaCl_2$ in the original mixture

<u>Example 22(b) A</u> spoon of mass 1.000-g is known to contain some silver. It was dissolved in HNO_3 and treated with Cl^-, 0.275 g of AgCl was obtained. What was the percentage by mass of silver in the ore?

$$Ag^+ + Cl^- \rightarrow AgCl$$

All the silver in the spoon is converted to AgCl by dissolving the spoon in HNO_3 and then treating in with Cl^-

Step 1: Calculate the number of moles AgCl and hence the number of moles of Ag

$$n_{Ag} = n_{AgCl}$$

$$n_{AgCl} = \frac{0.275\text{ g}}{143\text{ g mol}^{-1}}$$

$$= 0.001923\text{ mol}$$

Step 2: Calculate the mass of Ag

$$g_{ag} = n_{ag} \text{ X } MM_{ag}$$

$$= 0.001923\text{ mol x } 107.8\text{ g mol}^{-1}$$

$$= 0.2073\text{g}$$

Step 3: Calculate the % Ag in the sample

$$\%\text{ Ag} = \frac{g_{Ag}}{g_{sample}} \text{ X } 100$$

$$= \frac{0.2073\text{g}}{1.000\text{g}} \text{ x } 100 = 20.73\ \%$$

EXERCISE 16

16.1 A 0.833-g sample of a mixture of NaCl and $CaCl_2$ was dissolved in water and treated with Na_2CO_3 to precipitate $CaCO_3$. The precipitate was filtered and dried and found to weigh 0.415 g. What percentage by mass of the original sample was $CaCl_2$?

16.2 In a chemical analysis a 3.14-g sample known to contain $CuSO_4$ and $CUCl_2$ was dissolved in water and treated with $Ba(NO_3)_2$, Solid $BaSO_4$ was formed, which was filtered from the solution, dried and weighed. The $BaSO_4$ weighed 2.58 g.

(a) What mass Of SO_4^{2-} was in the $BaSO_4$?

(b) What was the weight percent SO_4^{2-} in the original sample?

16.3 A 0.400-g sample of an alloy of iron and nickel was dissolved in HCl to give Fe^{2+} and Ni^{2+}. The resulting solution was titrated with 22.3 mL of 0.0400 M $KMnO_4$, causing Fe^{2+} to be oxidized to Fe^{3+}. What percentage by mass of the alloy is iron? The net ionic equation for the reaction in the titration is

$$\mathbf{5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O}$$

16.4 When a solution containing copper(II) ton Is added to a solution containing I^- ion, the following reaction occurs.

$$\mathbf{2Cu^{2+} + 5I^- \rightarrow 2CuI + I_3}$$

In a chemical analysis of a copper ore, a sample weighing 0.2314 g was dissolved in acid and excess NaI solution was added. The I_3^- that was formed was titrated with 0.02000 M $Na_2S_2O_3$ solution. The titration required 22.94 mL of the $Na_2S_2O_3$ solution. What was the percentage by mass of copper in the ore? The reaction of I^{3-} with $S_2O_3^{2-}$ is as follows:

$$\mathbf{I_3^- + S_2O_3^{2-} \rightarrow 3I^{--} + S_4O_6{}^{2}}$$

Burette. A graduated glass tube fitted with a valve (stopcock) at one end. It Is used to dispense measured volumes of solutions.

Chemical analysis. The experimental determination of the composition of a substance

Endpoint. That point in a titration when delivery of the titrant is halted because an indicator changes color or some other event signals the completion of the reaction.

Standardizing a solution. Measuring accurately the concentration of solute in a solution. A standard solution is one whose solute concentration is accurately known.

Burrette Tap: A valve (for example, at the end of a burette) that is used to control the flow of a liquid or sometimes a gas.

Titrant : The solution dispensed from a burette

Titration An analytical procedure in which a solution, generally of known concentration, is added gradually from a burette to another solution where the solutes react. When the completion of the reaction is signaled by the change in the colour of an indicator the volume of the solution added from the burette is read and recorded.

Volumetric analysis: An analytical procedure that makes use of reactions in solution, where volumes and concentrations of solutions are carefully measured.

CONCENTRATION EXPRESSED AS MASS PERCENT

Another way of expressing the concentration of a solution is a ratio os the mass of solute to mass of solution expressed as a pecentage.

Mass percent thus refers to the mass of solute per 100g of solution.

Notation : eg 40% NaOH (m/m)

(m/m) = mass by mass

i.e. 40g of NaOH per 100g of solution

Example 23: A solution is prepared by dissolving 10g of NaOH in 150g of H_2O. Express the concentration of the solution in terms of mass percent.

All that is required is the ratio of the mass of solute to mass of solution multiplied by 100

$$\text{Mass \% (soln)} = \frac{\text{g solute}}{\text{g solution}} \times 100$$

$$= \frac{10g}{(10g + 150g)} x\,100$$

$= 6.25\%$ (m/m)

6.25% is the amount of solute in the solution.

Example 24(a) : How many grams of NaOH are present in 300g of 40% (m/m) NaOH solution

$$\text{Mass\% (soln)} = \frac{\text{g solute}}{\text{g soln}} \times 100$$

$$\text{g (solute)} = \frac{40 \times 300\text{g}}{100}$$

$= 120$g

<u>Exercise 24(b)</u> :.What volume of concentrated H_2SO_4 solution, which is 96 % H_2SO_4 by mass and has a density of 1.84 g mL^{-1} is required to prepare 400 mL of a 3.00 M is required to prepare 400 mL of a 3.00 M H_2SO_4 solution?

Assume 100 g of solution

$\therefore \frac{96}{100}$ x 100g = 96.0 g of H_2SO_4 (solute)

$$n(H_2SO_4) = \frac{g}{MM}$$

$$n(H\ SO_4) = \frac{96{,}0g}{98{,}0g\ mol^{-1}}$$

= 0.980 mol

Volume of 100 g of solution

$$V = \frac{g}{\rho} = \frac{100g}{1.84gm\ L^{-1}} = 54.3\ mL = 0.0543\ L$$

$$M_{H2SO4} = \frac{n}{V}$$

$$= \frac{0.980mol}{0.0543L}$$

$$= 18.0\ mol\ L^{-1}$$

18 M x y mL = 3M x 400 mL

y = 66.67 mL to 400 mL with water (vol = 400 mL – 66.67 mL)

<u>EXERCISE 17</u>

Add in more examples of percent composition

24. Concentrated HCl is 36.0% HCl by mass and has a density of 1.18 g mL.
    a) What is the molarity of the HCl?
    b) What volume of the concentrated acid is required to produce 15.0 L 0.346 mol L^{-1} HCl?

The Stoichiometry of Reactions in Solution using mass percent as the concentration unit

You must know how to use mass percent in dealing with the stoichiometry of reactions that take place in solutions. Most of the considerations that apply to molarity applies to mass percent

Example 25 : What mass of a 20% (m/m) aqueous solution of H_2SO_4 would be required to completely react with 3.00 g of zinc? The equation for the reaction

$Zn + H_2SO_4$ ® $ZnSO_4 + H_2$

Molar masses (in g mol^{-1}): Zn = 65.4 H_2SO_4 = 98.1

Determine moles of H_2SO_4 that reacts with Zn

$$n(H_2SO_4) = n(Zn)$$

$$= \frac{3.00\text{g}}{65.4\text{g mol}^{-1}}$$

$$= 0.0459 \text{ mol}$$

Determine g(H_2SO_4)

$$= 0.0459\text{mol} \times 98.1\text{g mol}^{-1}$$

$$= 4.5\text{g of } H_2SO_4$$

i.e. 4.5g of solute

Determine mass of solution

$$\text{Mass\% (solution)} = \frac{\text{g solute}}{\text{g solution}} \text{ x } 100$$

$$\therefore \text{g solution} = \frac{\text{g solute}}{\text{mass\%(soln)}} \text{ x } 100$$

$$= \frac{4{,}5g}{20.0} x\, 100$$

= 22.5g of H_2SO_4 solution

Note: Only the solute in the solution reacts.

<u>Example 26</u> : What volume of a 10% HNO_3 (m/m) solution. which has a density of 1.05g mL^{-1} will be required to react completely with 15.0g of $Ba(OH)_2$?

[molar masses (in g mol^{-1}) $Ba(OH)_2$ = 171 ; HNO_3 = 63.0]

$2\ HNO_3 + Ba(OH)_2$ ® $Ba(NO_3)_2 + 2H_2O$

Determine n**HNO_3**

$$\frac{nHNO_3}{2} = nBa(OH)_2$$

$nHNO_3 = 2 \times nBal(OH)_2$

$$= \quad 2 \times \frac{15{,}0g}{171g\ mol^{-1}}$$

= 0.175mol

Determine $gHNO_3$

= 0.175 mol x 63.0 g mol^{-1}

= 11.1 g

= grams of solute (i.e. pure HNO_3 only)

Determine The grams of HNO_3soln

$$\text{Mass\%(soln)} = \frac{g\ solute}{mass\%(so\ln)} \times 100$$

$$= \quad \frac{11.1g}{10} \times 100$$

= 111g of soln (HNO_3 + solvent)

Determine the volume of nitric acid solution.

Density $\quad \rho(soln) = \dfrac{g\ so\ln}{Vso\ln}$

$$Vsoln = \frac{g\ so\ln}{\rho\ so\ln} \quad = \quad \frac{111g}{1.05g\ mL^{-1}}$$

= 105.7 mL

EXERCISE 18

18.1. For the reaction:

$$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$$

What mass of $Ca(OH)_2$ is required to react completely with

a) 485 mL of 0.886 mol L^{-1} HCl?

b) 465 mL of an HCl solution that is 30,12% HCl, by mass, and has a density of 1.15g L^{-1}

18.2 Chlorine gas can be produced in the laboratory by the reaction indicated. A 61.3-g sample that is 96% $K_2Cr_2O_7$ is allowed to react with 320 ml of a hydrochloric acid solution having a density of 1.15 g mL^{-1} and containing 30% HCl, by mass. How much Cl_2 is produced, expressed in grams?

$$K_2Cr_2O_7 + 14\ HCl \rightarrow 2\ KCl + 2\ CrCl_3 + 7\ H_2O + 3\ C1_2(g)$$

18.3 A method employed to produce sodium sulfate, a substance used extensively in the textile industry, involves heating a mixture of ordinary salt and sulfuric acid.

$$2\ NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCl(g)$$

The sulfuric acid is in the form of a concentrated aqueous solution having a density of 1.73 g mL^{-1} and containing 80% H_2SO_4, by mass. How many litres of the sulfuric acid solution must be used to complete the reaction of 1.00 X 10^3 kg of salt?

SOLUTION CHEMISTRY: PAST EXAMINATION PAPERS

QUESTION 1 (5 MARKS)

A 1.00-g sample of an iron ore containing Fe_2O_3 was dissolved in acid and all the iron was converted to Fe^{2+}. The solution was titrated with 55.8ml of 0.0500 mol L^{-1} $KMnO_4$. The reaction that occurs is as follows

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$$

Molar masses (g mol^{-1}): Fe = 55.85, Fe_2O_3 = 159.7

(a) Calculate the mass of Fe^{2+} reacting **(3)**

(b) Calculate the mass percent of iron in the sample. **(2)**

QUESTION 2 (5 MARKS)

Oxalic acid ($H_2C_2O_4$) is present in many plants and vegetables. A vegetable-sample of mass 0.540 g requires 24.0mL of 0.0100 M $KMnO_4$ solution for complete reaction with the oxalic in the sample. The net ionic equation for the reaction is

$$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} ® 2Mn^{2+} + 10CO_2 + 8H_2O$$

Calculate the percent by mass of $H_2C_2O_4$ in the sample?

Molar mass of $H_2C_2O_4$ is 90.0 g mol^{-1}

QUESTION 3: 5 MARKS

A small piece of zinc is dissolved in 50,00 mL of 1,050 mol L^{-1} HCl. solution.

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

On completion of the reaction, the concentration of the 50,00 mL sample is redetermined and found to contain 0,766 mol of HCl per litre of solution.

a. Calculate the number of moles of HCl that reacted

b. Calculate the mass of the piece of zinc.

QUESTION 4: 10 MARKS

2.3 Calcium carbonate reacts with hydrochloric acid to produce calcium chloride, carbon dioxide and water:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

A 45,0-g sample of $CaCO_3(s)$ is added to 1,25 L of 7,87 mol L^{-1} HCl(aq) solution

2.3.1 Calculate the mass of excess reagent remaining after the reaction is complete.

2.3.2 Determine the molarity of the $CaCl_2$ solution after the reaction is complete.

Molar mass of $CaCO_3$ = 100 g mol^{-1}. (14)

QUESTION 5 : 8 MARKS

It is desired to prepare 20 L of an approximately 0,25 mol L^{-1} HCl solution, froma concentrated HCl solution, for use in a titration reaction in the laboratory.

a Calculate the molarity of concentrated hydrochloric acid that is 38% HCl by mass and has a density = 1,19 g mL^{-1}.

b What volume of the concentrated HCl acid (in 2.4.1) must be diluted to 20 L with distilled water to prepare a 0,25 mol L^{-1} (an approximate concentration) solution of the acid?

c A 25,00–mL sample of the approximately 0,25 mol L^{-1} HCl solution, prepared in (2.4.2) required exactly 30,10 mL of 0,2000 mol L^{-1} NaOH solution for complete neutralization. Calculate the exact molarity of the dilute HCl solution?

Molar mass of HCl = 36,46 g mol^{-1}

QUESTION 6 (14 MARKS)

A 30% HNO_3 (m/m) solution which has a density of 1.05g mL^{-1} will be required to react completely with 20.0g of $Ba(OH)_2$?

$$2HNO_3 + Ba(OH)_2 \rightarrow Ba(NO_3)_2 + 2H_2O$$

Molar masses(in g mol^{-1}): HNO_3 = 63.0 Ba $(OH)_2$ =171 $gmol^{-1}$

a) Calculate the mass of HNO_3 solute that will react with the given amount of $Ba(OH)_2$

b) Calculate volume, in litres, of the nitric acid solution that would be required for complete reaction

GASES

The bulk properties of a gas can be described in terms of the variables pressure (P), Volume (V) temperature (T) and Moles (n) If any three of the above properties are known the fourth one can be calculated.

Note: Pressure is defined as force per unit area

Units of Pressure

You must be familiar with the various ways of expressing pressure

SI Units	$N\ m^{-2} \equiv$ Pa
Other Units : =	1) mm of Hg $\equiv$ torr
	2) atmospheres (atm)

1 atm = 760mmHg = 760 torr = 1bar = 101325 Pa

Units of Temperature

Kelvin (K) - absolute temperature scale and Centigrade (^oC)

Relationship between K and ^oC

$K = {}^oC + 273$

Units of Volume

mL (cm^3), L(dm^3), M^3 (S.I)

GAS LAWS

The gas laws that will encountered in this course will include the following

1. Boyle's law
2. Charles' law
3. Combined Gas law
4. Avogadro's law
5. Ideal gas law
6. Dalton's law of partial pressure.

Boyles Law : For a given mass of gas maintained at constant temperature, volume is inversely proportional to pressure.

The relationship for Boyle's law expressed mathematically is

$$V \alpha \frac{1}{P}$$

or expressed as an equality

$$V = k\frac{1}{p}$$ where k is a propotionality constant

or $PV = K$ which is the equation for Boyle's law

or as an application equation for Boyles law

$$\mathbf{P_1 V_1 = P_2 V_2}$$

Charles' Law : For a given mass of gas maintained at constant pressure, volume is directly proportional to absolute temperatue

The relationship for Charles' law expressed mathematically is

$$V \alpha T$$

or expressed as an equality

$$V = kT$$ where k is a propotionality constant

or $\frac{V}{T} = K$ which is the equation for Charles' law

or as an application equation for Charles' law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Combined Gas Law (Boyles and Charles)

This equation is useful to convert volume of gas to moles of gas (see later).

Remember that

$V \alpha \frac{1}{P}$ (Boyle's Law) and that $V \alpha T$ (Charles')

Therefore the relationship for the combined gas law is

$V \alpha \frac{1}{P} T$ or expressed as an equation

$V = k\frac{T}{P}$ or $\frac{VP}{T} = k$

Or as an application equation of the combined gas law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This equation can be used to calculate the moles of a gas. (see Avogradros Law).

Avogadros Law: At constant temperature and pressure the volume of a gas is directly proportional to the number of moles of gas.

The relationship for Avogadro's law expressed mathematically is

$V \alpha n$

or expressed as an equality

$V = kn$ where k is a propotionality constant

or $\frac{V}{n} = k$ which is the equation for Avogadro's law

or as an application equation for Avogadro's law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Arising out of Avogadro's law is the fact that one mole of an ideal gas occupies a volume 22,414 L at STP which refers to **STANDARD TEMPERATURE** (O^oC or 273K) and STANDARD PRESSURE (1 atm or 101 325 Pa) i.e. constant T and P.

Ideal Gas Law

A summary of the three gas law described above

$V \alpha \frac{1}{P}$ (Boyles law - n, T constant)

$V \alpha T$ (Charles law - n, P costant)

$V \alpha n$ (Avogadros law - P,T const)

Combining the three gas laws above gives the relationship

$$\therefore\ V\ \alpha \frac{nT}{P}$$

or $V = \frac{RnT}{P}$ expresses an equality which when rearranged gives rise to

$$PV = nRT$$

where R is the ideal gas constant. This equation is called the ideal gas equation. It shows the relationship between the four variables (P, T, V and n) of a gas. From the Chemist's point of view the equation is most useful in the form

$$\therefore n_{gas} = \frac{PV}{RT}$$

where the number of moles of a gas can be obtained from the variables P, T, V that can be measured in the laboratory.
Remember that this equation applies only to gases. You cannot calculate the moles of a liquid or solid using this equation.

<u>Value and Units of R</u>

1. In terms of S.I. units

$$R = \frac{PV}{nT}$$

For 1 mol of ideal gas at STP

$$R = \frac{101325\,\text{Nm}^{-1}\ \text{x}\ 0{,}022414\text{m}^{3}}{1\text{mol x } 273{,}15\text{K}}$$

$= 8{,}314\ \text{N.m mol}^{-1}\ \text{k}^{-1}$

$= 8{,}314\ \text{J mol}^{-1}\ \text{k}^{-1}$

Note : Pressure must be expressed in pascals (Nm^{-2}) and volume must be expressed in m^3 ($m^3 \equiv dm^3 \times 10^{-3}$)

2. In terms of the unit of atms and litres

$$R = \frac{1\text{atm x } 22{,}414\ \text{L}\ (\text{dm}^{3})}{1\text{mol x } 273{,}15\text{k}}$$

$= 0{,}0820\ \text{atm. L mol}^{-1}\text{K}^{-1}$

Gas density and Molar mass of a gas

$$n(g) = \frac{PV}{RT}$$

or $$\frac{g(g)}{MM} = \frac{PV}{RT}$$

rearrange

$$\frac{g(gas)}{V(gas)} = \frac{P \text{ x } MM}{RT}$$

But $$\frac{g}{V} = P(density)$$

$\therefore$ P(gas) = $$\frac{P \text{ x } MM}{RT}$$

rearranging this equation

MM(gas) = $$P\frac{RT}{P}$$

The molar mass of an unknown gas can be determined by measuring its P, T and P, in the laboratory, and this information can be used in to determine the formula of a compound – refer to empirical formula calculations

Example : A sample of N_2 gas that has a volume of 30.0L exerts a pressure of 0,987 atm at 30^oC. Determine the number of moles of N_2 without using the ideal gas equation.

1. Determine the vol of N_2 at STP

STP EXP

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_1 = \frac{P_2V_2T_1}{P_1T_2}$$

$$= \frac{0.987\text{atm x } 30.0\text{L x } 273\text{k}}{1\text{atm x } 303\text{k}}$$

$=$ 26.7 L of N_2 at STP

(Remember : 1 mol of gas at STP occupies a volume of 22.4L)

$$\therefore n_{N2} = \frac{V(N_2\ )\text{ at STP}}{22.4\text{ L mol}^{-1}}$$

$$= \frac{26.7\text{L}}{22.4\text{L mol}^{-1}}$$

$=$ 1,19 mol of N_2

Example : The density of an unknown gas, at STP, is 1,429g L^{-1}. Calculate the molar mass of the gas. (R = 0,0821 atm.L. $mol^{-1}K^{-1}$)

$$MM = P\frac{\text{x R x T}}{P}$$

$$= \frac{1.429\text{gL}^{-1}\text{ x } 0.0821\text{atm L.mol}^{-1}\text{K}^{-1}\text{ x } 273\text{K}}{1\text{atm}}$$

$=$ 32.02 g mol^{-1}

Example : A 0.100g sample of a compound of empirical formula CH_2F_2 occupies 0.0470 L at 298k and 755 mm of Hg

(a) What is the molar mass of the compound?

(b) What is the molecular formula of the compound?

(R as per previous example)

$$Patm = \frac{755\text{ mm Hg}}{760\text{ mmHg atm}^{-1}} = 0.993\text{ atm}$$

$$\frac{g}{MM} = \frac{PV}{RT}$$

$$MM = \frac{g \times R \times T}{PV}$$

$$= \frac{0.100g \times 0.0821 atm\ L\ mol^{-1}k^{-1} \times 298k}{0.993 atm \times 0.0470L}$$

= 52.4 g mol^{-1} which is the molar mass of CH_2F_2

EMF of CH_2F_2 = 52.0g mol^{-1}

Since MM = EFM

Molecular formula is the same as the empirical formula.

Gas Mixtures and Partial Pressures

In a mixture of gasses the total number of moles of gases is taken as the sum of the moles of the component gases

$n_{Total} = n_A + N_B + \ldots..$

e.g. A mixture of 0,5 mole O_2 and 0,5 mol N_2 has the same effect as 1 mole of either gas

$PV = n_{Total} RT$

$= (n_{O2} + n_{N2})RT$

$$P_{Total} = \frac{n_{O_2}RT}{V} + \frac{n_{N_2}RT}{V}$$

$P_{Total} = P_{O2} + P_{N2}$

Which is the mathematical expression for Dalton's Law of Partial pressure which states that the total pressure of a mixture of gases, which do not react, is equal to the sum of the partial pressures of the individual gases

i.e. $P_T = P_A + P_B + P_C + \ldots.$

Application to a chemical system

Gases are often collected over water and the pressure exerted by the "wet" gas is due to the gas as well as water vapour.

i.e. $P_{total} = Pg + P_{H2O}$

$P_{g(dry)} = Pt - P_{H2O}$

The vapour pressure of H_2O varies with temperature. It has been determined over a range of temperatures and these values have been tabulated.

Example : Oxygen is collected <u>over water</u> at 25°C in a 2.0 L vessel at a total pressure of 765 mmHg. Calculate the number of moles of O_2 produced. The vapour pressure of H_2O at 25°C = 25 mm of Hg

$R = 0.0821$ atm.L.mol^{-1}k^{-1}

Determine the pressure of dry O_2 gas

$P_{O2} = P_T - P_{H2O}$

$= (765 – 25)$ mm Hg

$= 740$ mm Hg

Convert mmHg to atm

$$= \frac{740 \text{ mmHg}}{760 \cdots \text{mmHg atm}^{-1}}$$

$= 0,9737$ atm

$$n_{O2} = \frac{PV}{RT}$$

$$= \frac{0.9737 \text{ atm x } 2.0 \text{ L}}{0.0821 \text{ atm.L.mol}^{-1}\text{k}^{-1} \text{ x } 298\text{k}}$$

$= 0.0796$ mol

As an exercise repeat same calculation using SI units.

<u>STOICHIOMETRY OF GASES</u>

Example : Calculate the mass of MnO_2 needed to produce 4,00L of Cl_2 gas, at STP, according to the equation

$MnO_2 + 4HCl$ ® $MnCl_2 + 2H_2O + Cl_2$

MM of MnO_2 = 86,9g mol^{-1}

$n_{MnO2} = n_{Cl2}$ (from equation)

Example: In an experiment to determine the atomic weight of Al, 1.349 of Al metal was allowed to react with excess dilute H_2SO_4 and 1.910 L of H_2 gas was evolved and collected over water at 23°C and 746 mm of Hg. Calculate the atomic weight of Al.

Vapour pressure of H_2O at 23°C = 21.0 mmHg (Two methods).

METHOD 1

Determine the partial pressure of H_2

$P_{H2} = P_T - P_{H2O}$

$= (746 – 21)$ mmHg

$= 725$ mm Hg

$$P_{H2} \text{ in atm} = \frac{725 \text{ mmHg}}{760 \text{ mmHg atm}^{-1}}$$

$= 0.954$ atm

Determine number of moles of H_2

$$n_{H2} = \frac{PV}{RT}$$

$$= \frac{0.954 \text{ atm x } 1.910\text{L}}{0.0821 \text{ atm.L.mol}^{-1}\text{k}^{-1} \text{ x } (23 + 273)\text{k}}$$

= 0.0750 mol H_2 gas

Determine number of moles of aluminium

$$\frac{n_{Al}}{2} = \frac{n_{H_2}}{3} \text{ (from the balanced equation)}$$

$$n_{Al} = \frac{2}{3} \text{ x } n_{H2}$$

$$= \frac{2}{3} \text{ x } 0{,}0750 \text{ mol}$$

= 0,0500 mol of Al

Determine atomic weight of Al

$$MM_{Al} = \frac{g_{Al}}{n_{Al}}$$

$$= \frac{1.349\text{g}}{0.0500\text{ mol}}$$

$$= 26.98\text{ g mol}^{-1}$$

On atomic scale = 26,98 u

METHOD 2

Determine V of H_2 at STP conditions using the general gas law (combined gas law)

STP conditions Experimental conditions

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$VH_2 = \frac{P_2V_2 \text{ x } T_1}{T_2 \text{ x } P_1}$$

$$= \frac{725\text{ mmHg x } 1.910\text{L x } 273\text{K}}{296\text{K x } 760\text{ mmHg}}$$

$$\therefore V_{H2}\text{ (at STP)} = 1.680\text{ L}$$

$$nH_2 = \frac{V_{H_2}\text{ at STP}}{22.4\text{ L mol}^{-1}}$$

$$= \frac{1.680\text{ L}}{22.4\text{ L mol}^{-1}}$$

$$= \mathbf{0{,}075\ mol}$$

Determine the number of moles and atomic weight of aluminium by in the same way as for method 1.

GASES: PAST EXAMINATION PAPERS

QUESTION 1 (7 MARKS)

A 5.65g sample of Al is reacted with excess of HCl (aq) and the liberated H_2 (g) is collected over water at 26^{0}C and a pressure of 746 mmHg. The equation for the reaction is

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

Molar masses (gmol^{-1}) Al = 26.98, HCl = 36.40

Vapour pressure of H_2O at 26^{0}C = 25.2 mmHg

R= 8.314 J mol^{-1} K^{-1} R = 0.0821 atm L mol^{-1} K^{-1}

(a) Calculate the number of moles of H_2 (g) liberated. **(3)**

(b) What volume of gas, in mL, is collected? **(4)**

QUESTION 2 (4 MARKS)

Automotive air bags inflate when sodium azide, NaN_3, rapidly decomposes into its component elements:

$$\mathbf{2\ NaN_3 \rightarrow 2\ Na\ (s) + 3\ N_2\ (g)}$$

Calculate the mass of sodium azide, in grams, that are required to produce 283 000 mL of nitrogen gas if the gas has a density of 1.25 grams per litre?

Atom masses: N = 14.0 g mol^{-1} Na = 23.0 g mol^{-1}

QUESTION 3 : 6 MARKS

2.1 Analysis of a gas shows that it contains carbon (37,23% by mass), hydrogen (7,81% by mass) and chlorine. At 150^{0}C and 1,00 atm pressure, 500 mL of the gas has a mass of 0,922 g.

a Determine the empirical formula of the gas.

b Calculate the molar mass of the gas

c Determine the molecular formula of the gas.

Molar masses(in g mol^{-1}): H = 1,01 C = 12,0 Cl = 35,5

R = 0,0821 atm L mol^{-1} K^{-1}

QUESTION 4 (5 MARKS)

Gaseous ammonia is synthesized from nitrogen and hydrogen by the reaction

$$\mathbf{N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)}$$

When a 355L sample of H_2(g), at 25.0^{o}C and 542mmHg, is combined with excess N_2 gas a reaction occurs in which 2.5 g ammonia gas is produced.

Calculate the percentage yield of NH_3(g) for this reaction

Molar masses (in g mol^{-1}): H = 1.01 N = 14.0

R = 0.0821atm.L.mol^{-1} K^{-1}

EXAMPLES: STOICHIOMETRY AND RELATED CALCULATIONS DISCUSSED AT LECTURES

Example 1 Calculate the atomic weight of chlorine, which consists of 75.77% ^{35}Cl (34.97amu) and 24.23% ^{17}Cl (36.95 amu).

Example 2: How many moles are there in 125 g of Al (MM of Al = 27.0g mol^{-1})

Example 3 : How many grams are there in 5.00 mol of NaOH
(MM of NaOH = 40.0g mol^{-1})

Example 4 : How many carbon atoms are there in a 1 carat diamond, which is pure carbon 1 carat has a mass of exactly 0,2000g.

Example 5 : Calculate the percentage composition of H_3PO_4
[molar masses (in g mol^{-1}) H = 1,01 P = 31.0 O = 16,0

Example 6.

A 4.00-g sample of a copper-bromine compound was decomposed, yielding 1.14 g of pure copper. What is the empirical formula of the compound?

Example 7: Carotene has a percentage composition of 89.49% C and 10.51% H. Its molecular mass was found to be 536.8 u
Calculate its (i) empirical formula (ii) Molecular formula

Example 8: Vitamin C is a compound that contains the elements C. H and O. Complete combustion of a sample of mass 0.2000 g of vitamin C produced 0.2998 g of CO_2 and 0.08185 g of H_2O. Determine the empirical formula of vitamin C. Molar masses (in g mol^{-1}) : CO_2 = 44.01 H_2O = 18.02
C = 12.01 H = 1.008 O = 16.00

<u>Example 9;</u> The compound caffeine contains the elements C. H and O. Combustion analysis of a 1.500g sample of caffeine produces 2.737g of CO_2 and 0.6814g of H_2O. Further analysis of another sample, of mass 2,500g of caffeine produces 0,8677g of NH_3. Determine the empirical formula of caffeine.

<u>Example 10</u>: How many moles of $H_2C_2O_4$ (oxalic acid) are required to react completely with 1.50 mole of $KMnO_4$ according to the following equation

$2KMnO_4 + 6HCl + 5H_2C_2O_4$ ® $2MnCl_2 + 10CO_2 + 2KCl + 8H_2O$

<u>Example 11</u>: How many moles of Al_2O_3 are produced from 81.0 g of Al. The equation for the reaction is

$Fe_2O_3 + Al$ ® $Al_2O_3 + Fe$

(MM of Al = 27.0g mol^{-1})

<u>Example 12</u> : The decomposition of $KClO_3$ occurs as shown below:

$2KClO_3$ ® $2KCl + 3O_2$

Calculate the number of grams of $KClO_3$ required to produce 4.50 mol of O_2

(MM of $KClO_3$ = 122.6g mol^{-1})

<u>Example 13:</u> Calculate. the number of grams of $MgCl_2$ that could be obtained from 8.50 g of HCl when the latter is reacted with excess MgO. MM (in g mol^{-1}) : HCl = 36.5 $MgCl_2$ = 95.3

<u>Example 14</u> Zinc and oxygen combine to produce zinc oxide according to the equation

$$2Zn + O_2 \rightarrow 2ZnO$$

How many grams of ZnO will be formed if 14.3 g of Zn are mixed with 3.72 g of O_2?

<u>Example 15</u>: A 50,0g sample of $CaCO_3$ is reacted with 35,0g of H_3PO_4

$3CaCO_3 + 2H_3PO_4$ ® $Ca_3(PO_4)_2 + 3CO_2 + 3H_2O$

(a) Determine which reactant is the limiting reagent.

(b) Calculate the mass of $Ca_3(PO_4)_2$ formed when the reaction is complete.

(c) Calculate the mass (in grams) of excess reagent remaining after the reaction is complete

Molar masses (in g mol^{-1}): $Ca_3(PO_4)_2$ = 310,0g mol^{-1}

$CaCO_3$ = 100,1 H_3PO_4 = 98,03

<u>Example 16</u> : The gas diborane B_2H_6 can be prepared by the following reaction

$$\mathbf{3NaBH_4 + 4BF_3 ® 3NaBF_4 + 2B_2H_6}$$

if 18,9g of $NaBH_4$ and an excess of FB_3 is reacted and 7,50g of B_2H_6 is isolated. What is the percent yield of B_2H_6?

Molar masses (in g mol^{-1}): $NaBH_4$ = 37.8 B_2H_6 = 27.7

<u>Example 17:</u> A student prepared a solution of NaCl by dissolving 1,461g of NaCl in a 250 mL volumetric flask. What is the molarity of the solution?

Molar mass : NaCl = 58,44 g mol^{-1}.

<u>Example 18</u>: How many mL of 0,250 mol L^{-1} NaCl solution must be measured to obtain 0,100 mol of NaCl.

Examples 19

(a) What volume of a 18 mol L^{-1} H_2SO_4 solution is required to prepare 2.50 litres of a 0.360 mol L^{-1} solution of H_2SO_4?

(b) If 82.5 mL of 6.25 mol L^{-1} solution to a final volume of 250 mL. What is the molarity of the dilute HCl solution?

<u>Example 20:</u>Copper is dissolved in dilute HNO_3 solution by the following reaction

$$3Cu + 8HNO_3 ® Cu(NO_3)_2 + 2NO + 4H_2O$$

How many mL of 3,00 mol L^{-1} HNO_3 solution can react with 10 g of Cu?

Molar mass of Cu = 63,5 g mol^{-1}

<u>Example 21</u>: A soda mint tablet contains $NaHCO_3$ as an antacid. One tablet requires 34.5 mL of a 0.138 mol L^{-1} HCl solution for complete reaction. Calculate the number of grams of $NaHCO_3$ that one tablet contains. The equation for the reaction is given by

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Molar mass of $NaHCO_3$ = 84.0 g mol^{-1}.

<u>Example 22</u>

(a) How many milliliters of 0.200 M $Ba(NO_3)_2$ are needed to react completely with 75.0 mL of 0.150 M $Fe_2(SO_4)_3$ solution? The net ionic equation for the reaction is

$$\mathbf{Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4(S)}$$

(b) In a titration how many millilitres of a 0.150 mol L^{-1} $KMnO_4$ solution are required to react 15.0 mL of a 0.250 mol L^{-1} $FeCl_2$ solution. The equation for the reaction is:

$$KMnO_4 + FeCl_2 + 8HCl \rightarrow 5FeCl_3 + MnCl_2 + KCl + 4H_2O$$

Example 23 When a 1.000-g sample of an ore known to contain silver was dissolved in HNO_3 and treated with Cl^-, 0.275 g of AgCl was obtained. What was the percentage by mass of silver in the ore?

Example 24

A 0.4308 g sample of iron ore is dissolved in acid and the iron converted into the Fe^{2+}. This solution is reacted with a solution of $KMnO_4$. The reaction requires 27.35 mL of a 0.02469 mol L^{-1} solution of $KMnO_4$. Calculate the percentage of Fe in the ore. The equation for the reaction is:

$$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5\ Fe^{3+} + Mn^{2+} + 4H_2O$$

Molar mass of Fe = 55,85 g mol^{-1}

<u>Example 25</u>: A solution is prepared by dissolving 10 g of NaOH in 150 g of H_2O. Express the concentration of the solution in terms of mass percent.

Example 26 How many grams of NaOH are present in 300 g of 40% (m/m) NaOH solution

<u>Example 27</u> : What mass of a 20% (m/m) aqueous solution of H_2SO_4 would be required to completely react with 3.00 g of zinc? The equation for the reaction

$Zn + H_2SO_4$ ® $ZnSO_4 + H_2$

Molar masses (in g mol^{-1}): Zn = 65.4 H_2SO_4 = 98.1

<u>Example 28</u> : What volume of a 10% HNO_3 (m/m) solution. which has a density of 1.05g mL^{-1} will be required to react completely with 15.0g of $Ba(OH)_2$?

[molar masses (in g mol^{-1}) $Ba(OH)_2$ = 171 ; HNO_3 = 63.0]

$2 HNO_3 + Ba(OH)_2$ ® $Ba(NO_3)_2 + 2H_2O$

<u>Example 29</u> (a) A sample of nitrogen gas that has a volume of 30.0 L exerts a pressure of 0.987 atmospheres at 30^0C.

Determine the number of moles of nitrogen without using the ideal gas equation.

<u>Example 29 (b)</u> The density of an unknown gas, at STP, is 1.429 g L^{-1}. Calculate the mass of the gas. (R = 0.0821 atm.L.mol^{-1} K^{-1})

<u>Example 30</u> A 0.100 g sample of a compound of empirical formula CH_2F_2 occupies 0.0470 L at 298 K and 755 mm of Hg.

(a) What is the molar mass of the compound?

(b) What is the molecular formula of the compound?

R = 0.0821 atm.L.mol^{-1} K^{-1})

<u>Example 31</u> Calculate the mass of MnO_2 needed to produce 4.00 L of Cl_2 gas, at STP, according to the equation

$MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

Molar Mass of MnO_2 = 86,9 g mol^{-1}

<u>Example 32</u>

The decomposition of 13.14 g of $Sr(IO_3)_2$ results in the production of SrI_2 and 4.0321 L of O_2 measured at 136.5^{0}C and 570 mm Hg pressure. The SrI_2 thus produced is quantitatively converted into $SrCl_2$ of which 4.77 g is obtained.

Calculate the molar masses of **Sr** and **I** given that the Molar masses of Cl and O are 35.5 and 16.0 g mol^{-1}.

Example 33

Oxygen is collected over water at 25^0C in a 2.0 L vessel at a total pressure of 765 mm Hg. Calculate the number moles of O_2 produced. The vapour pressure of H_2O at 25^0C = 25 mm of Hg

Example 34

In an experiment to determine the atomic weight of Al, 1.349 g of Al metal was allowed to react with excess dilute H_2SO_4 and 1.910 L of H_2 gas was evolved and collected over water at 23^0C and 746 mm of Hg. Calculate the atomic weight of Al. Vapour pressure of H_2O at 23^0C and 21.0 mm of Hg.