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## SENIOR FIVE TERM 1

### TOPIC 1/3: Moles and Equations

#### Atomic Structure

An atom is the smallest electrically neutral particle that takes part in a chemical reaction.

An atom is composed of electrons, protons and neutrons the properties of which are given in the table below:

components	symbol	charge	mass
Proton	p	+1	1
Neutron	n	0	1
Electron	e	-1	$\frac{1}{1840}$

The proton and neutron are situated in the nucleus whereas the electrons rotate around the nucleus in definite shells.

#### Definitions

**Atomic number** is the number of protons present in the nucleus of an atom. All atoms of the same element have the same number of protons and thus the same atomic number.

**Atomic mass** is the sum of protons and neutron in an atom.

**Isotopes** are atoms of the same element having the same atomic number but different number of neutron and thus different atomic mass.

**Relative atomic mass** is the number of time atoms of an element are heavier than a  $^{12}\text{C}$  of carbon-12 isotope. Relative atomic mass is obtained as an average of atomic mass of isotopes of an element in the ratio of their existence or relative abundance.

#### Example 1

Natural chlorine has isotopes  $^{35}\text{Cl}$  and  $^{37}\text{Cl}$  the percentage of the abundance of these isotopes is 75% and 25% respectively. Calculate the atomic mass of Cl-atom.

#### Solution

The question implies that for every 100 atoms of chlorine, 75 of them each weigh 35 units and 25 of them each weigh 25 units.

Thus mass of 100 atoms of chlorine =  $75 \times 35 + 25 \times 37$

**Relative atomic mass of chlorine = average mass of 100 atoms** =  $\frac{35 \times 75 + 37 \times 25}{100} = 35.50$

### Trial 1

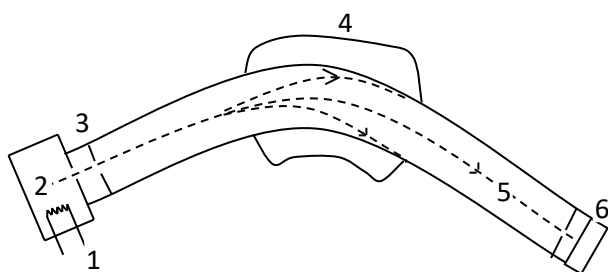
1. The relative atomic mass of copper is 63.5. Calculate the percentage of the isotopes  $^{63}\text{Cu}$  and  $^{65}\text{Cu}$  in it. [75%, 25%]
2. Naturally occurring boron consist of two isotopes whose atomic weights are 10.01 and 11.01. the atomic weight of boron is 10.81. Calculate the % of each isotope in naturally occurring boron [82.9%, 17.1%]
3. Naturally occurring carbon consist of two isotopes namely  $^{12}\text{C}$  and  $^{13}\text{C}$ . What is the % abundance of the isotopes in a sample of

carbon whose atomic weight is 12.01112? Assume that the nucleic masses of  $^{12}\text{C}$  and  $^{13}\text{C}$  are 12.0 and 13.00339 respectively [98.892%, 1.108%]

4. A sample of ordinary neon is found to consist of  $^{20}\text{Ne}$ ,  $^{21}\text{Ne}$  and  $^{22}\text{Ne}$  isotopes in percentage of 90.92%, 0.26% and 8.82% respectively. Calculate the actual atomic weight of neon. [20.179]

### Determination of Relative atomic masses

#### Mass spectrometer



#### Functions at each numbered step

1. Heated filament gives electrons. They pass into the ionization chamber.
2. The sample to be analyzed is injected as a gas into the ionization chamber. Electrons collide with and ionize the molecules of the sample.
3. To this plate a negative potential is applied (about 800V). The electric field accelerates the positive ions while repelling the negative ions.
4. An electromagnet produces a magnetic field. The field deflects the beam of ions

into circular paths. Ions with a high ratio of charge to mass are deflected more than those with low ratio of charge/mass.

5. By varying the magnetic field ions that have a correct charge/mass ratio are directed through the slit to the collector.
6. Amplifier. Here the charge received by the collector is turned into a sizable electric current, interpreted and recorded.
7. If the magnetic field is kept constant while the accelerating voltage is continuously varied, one species after another is deflected into the ion collector.

#### Uses of mass spectrophotometer

1. Determination of the relative molecular mass of a compound.
2. Identification of compounds
3. In forensic science it is useful because a small sample gives results.

## Trial 2

- 1 (a) Define the terms; mass number, isotope, relative atomic mass
- (b) Chlorine has two isotopes of relative atomic masses 34.97 and 36.96 and relative abundance 75.77% and 24.23% respectively.

Calculate the mean relative atomic mass of naturally occurring chlorine [35.452]

- 2 Element X with atomic number 84 and mass number 216 decays by loss of an  $\alpha$ -particle gives element Y. Y further decays by loss of  $\beta$ -particle to give Z.

- 3 (a) write down
- (i) the mass number of element Y [82]
- (ii) the atomic number of element Z [83]
- (b) In which group of the Periodic Table would you expect Z? [5]
- (c) The table shows the mass numbers and the percentage abundance of an element.

Mass number	% abundance
54	5.84
56	91.68
57	2.17
58	0.31

Calculate the relative atomic mass of Q [55.9111]

- 4 (a) (i) Define the term isotope
- (ii) Describe the main steps involved in the operation of a mass spectrometer (diagram not required)
- (b) The table below shows the information from a mass spectrum of a lead sample.

Isotope	Detector/mA
---------	-------------

204	0.16
206	2.72
207	2.50
208	5.92

Calculate

- (i) The relative abundance of the different isotopes of lead in the sample used [1.416%, 24.071%, 22.12%, 52.389%]
- (ii) The relative atomic mass of lead [207.24]
- (iii) State two advantages of using a mass spectrometer over depression of freezing point method of determining relative atomic mass.
- (c) The initial count of a radioactive nucleus was  $680\text{s}^{-1}$ , after 350s, the count rate was  $125\text{s}^{-1}$ . Calculate
- (i) decay constant
- (ii) half-life of the nucleus

- 5 The two isotopes of hydrogen having mass number 1 and 2 have atomic weight equal to 1.007825 and 2.01473 as determined by mass spectrograph. These isotopes are present in ordinary hydrogen in ratio of 6400 to 1. Determine the mean-atomic weight of hydrogen atoms. [1.00798]

- 6 The mass spectrum of chlorine shows peaks at masses 70, 72, and 74. the heights of the peaks are in ratio of 9:6:1

Calculate

- a. The average atomic mass of chlorine (2marks) [35.5]
- b. The relative abundance of  $^{35}\text{Cl}$  and  $^{37}\text{Cl}$ . (3 marks) [75%, 25%]

- 7 (a) Define the term relative atomic mass (2mark)
- b. Explain how the relative atomic mass can be determined by the mass spectrometer (10 marks)
- c. The mass spectrum of an element, A, contained four lines at mass/charge of 54, 56, 57, and 58 with relative intensities of 5.84, 91.68, 2.17 and 0.31 respectively
- Explain what the term relative intensities means
  - Calculate the relative atomic mass of A [55.9111]

- 8 An element Y has three naturally occurring isotopes with isotopic masses and relative abundances' as shown below:.

Isotopic mass	Relative abundance (%)
23.98	78.60
24.98	10.11
25.98	11.29

Calculate the relative atomic mass of Y

(2 marks)[24.31]

### Formula masses (relative formula mass or Molecular mass)

This is the number of times a molecule is heavier than one twelfth of a carbon-12 atom. It is calculated by addition of the individual RAM of the constituent atoms in a molecules, Relative formula mass has no units

#### Example 2

What is the formula mass of Ozone,  $O_3$ ?

#### Solution

The Ozone molecule contains three oxygen atoms, each of which weighs 16 grams.

$$\therefore \text{the formula mass of Ozone is } 3 \times 16 = 48$$

#### Example 3

What is the formula mass of Ammonia,  $NH_3$ ?

#### Solution

RAM of nitrogen atom is 14 and that of hydrogen atom is 1. There are three

hydrogen atoms in the molecule and one nitrogen atom;

$$\therefore \text{the formula mass of ammonia is: } 14 + (3 \times 1) = 17 \text{ grams.}$$

#### Example 4

What is the formula mass of Glucose,  $C_6H_{12}O_6$  ?

(C= 12, H=1, and O=16)

#### Solution

$$\begin{aligned} \therefore \text{The formula mass of Glucose is} \\ &= 6 \times 12 + (12 \times 1) + (6 \times 16) \\ &= 72 + 12 + 96 \\ &= 180 \end{aligned}$$

## Trial 2

1. Calculate the formula mass of

(a) Ammonium sulphate,  $(\text{NH}_4)_2\text{SO}_4$  [132]

(b) Sodium carbonate decahydrate,  
 $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . [286]

(c) Oxalic acid,  $\text{HOOC-COOH}$  [90]

(d) Maleic acid  $\text{HOOC-CH=CH-COOH}$  [116]

## The mole

A mole is the AMOUNT of a substance which contain as many elementary units (particles, i.e., molecules, atoms, ions etc) as there are in 12 g of carbon-12 isotope ( $^{12}_6\text{C}$ ). A mole of any substance contains  $6.02 \times 10^{23}$  particles.

The number of particles in any mole of a substance ( $6.02 \times 10^{23}$ ) is called **AVOGADRO'S NUMBER, L**.

**Molar Mass** is the mass in grams of 1 mole of a substance. It is numerically equal to its relative atomic mass or its relative formula mass. e.g., One mole of carbon weighs 12 g, 1 mole of oxygen molecule weighs 32 g, and 1 mole of ammonium sulphate weighs 132 g.

### Calculations using the Avogadro number

The Avogadro number, L, is taken as  $6.0 \times 10^{23}$  in the following examples.

#### Example 5

*How many C atoms are there in a carbon rod weighing 8 g (C=12).*

#### Solution

12 g (RAM of carbon in g or 1 mole) of carbon have  $6.0 \times 10^{23}$  atoms.

$$\begin{aligned} \therefore 8 \text{ g of carbon has } & \frac{8 \times 6.0 \times 10^{23}}{12} \text{ atoms} \\ & = 4.0 \times 10^{23} \text{ atoms} \end{aligned}$$

#### Example 6

How many grams of copper (Cu=64) contain  $4.5 \times 10^{23}$  atoms?

#### Solution

$6.0 \times 10^{23}$  atoms are in 1 mole of copper i.e. 64 g

$$\begin{aligned} \therefore 4.5 \times 10^{23} \text{ atoms are in } & \frac{4.5 \times 10^{23} \times 64}{6.0 \times 10^{23}} \\ & = 48 \text{ g} \end{aligned}$$

#### Example 7

*How many (i) molecules (ii) atoms are there in 35.5 g of gaseous chlorine  $\text{Cl}_2$ ?*

#### Solution

(i) Mass of 1 mole of chlorine molecule,  $\text{Cl}_2 = 71 \text{ g}$

i.e. 71 g of chlorine contain  $6.0 \times 10^{23}$  molecules.

$\therefore$  35.5 g of chlorine contains

$$\begin{aligned} & = \frac{35.5 \times 6.0 \times 10^{23}}{71} \\ & = 3.0 \times 10^{23} \text{ molecules} \end{aligned}$$

(ii) Each chlorine molecule contains 2 atoms. Hence there are

$2 \times 3.0 \times 10^{23}$  atoms in 35.5 g of gaseous chlorine,  $\text{Cl}_2$

∴ No. of atoms in 35.5 g of chlorine gas =  $6.0 \times 10^{23}$  atoms.

### Example 8

How many ions are there in 13.5 g of copper (II) chloride,  $\text{CuCl}_2$ ?

$$\begin{aligned}\text{Mass of 1 mole of CuCl}_2 &= 64 + (2 \times 35.5) \\ &= 135 \text{ g}\end{aligned}$$

There are 3 ions (one copper and two chloride ions) in  $\text{CuCl}_2$  molecule.

i.e. 1 mole or 135 g of  $\text{CuCl}_2$  contain  $3 \times 6.0 \times 10^{23}$  ions.

$$\begin{aligned}\therefore 13.5 \text{ g of CuCl}_2 \text{ contain } &\frac{13.5 \times 36.0 \times 10^{23}}{135} \\ &= 1.8 \times 10^{23} \text{ ions.}\end{aligned}$$

[NB. it contains  $0.6 \times 10^{23}$  copper ions and twice that i.e.  $1.2 \times 10^{23}$  chloride ions.]

### The 'mole' concept in calculations

To convert Mass of a substance into moles

$$\text{No. of moles} = \frac{\text{Mass of substance}}{\text{Mass of 1 mole}}$$

To convert moles of a substance in grams

Mass in grams = No. of moles  $\times$  mass of 1 mole.

### Example 9

How many moles of carbon dioxide molecules are present in 11 g of carbon dioxide,  $\text{CO}_2$ ?

$$\text{Mass of 1 mole of CO}_2 = 12 + (2 \times 16) = 44 \text{ g}$$

$$\begin{aligned}\text{No. moles of CO}_2 &= \frac{\text{Mass in grams}}{\text{Mass of 1 mole}} \\ &= \frac{11}{44} \text{ moles} = 0.25 \text{ moles}\end{aligned}$$

### Example 10

What is the mass of 3 mole of carbon monoxide?

$$\text{The Molar mass of CO} = (12 + 16) = 28 \text{ g}$$

Mass in grams = No. of moles  $\times$  mass of 1 mole.

$$= 3 \times 28$$

$$= 84 \text{ g}$$

### Example 11

What mass of magnesium ( $\text{Mg}=24$ ) would contain the same number of atoms as 4 g of carbon. ( $\text{C}=12$ )

Solution

No. of moles of carbon =

$$\frac{\text{Mass in grams}}{\text{Mass of 1 mole}} = \frac{4}{12} \text{ moles}$$

$$= 0.333 \text{ mol}$$

A mole of any element contains the same number of atoms.

Hence 0.333 moles of magnesium contain the same number of atoms as 4 g of carbon.

Mass of 0.333 moles magnesium

$$= \text{No. of moles} \times \text{Mass of 1 mole}$$

$$= 0.333 \times 24$$

$$= 8 \text{ g.}$$

### Example 12

Calculate the mass of

(i) 1 atom of magnesium.

(ii) 5 molecule of ethane.

### Solution

(i) Relative atomic mass of magnesium = 24

$$\begin{aligned} 1 \text{ atom of magnesium} &= \frac{24 \times 1}{6.0 \times 10^{23}} \\ &= 4 \times 10^{-23} \text{ g} \end{aligned}$$

(ii) Molar mass of  $\text{C}_2\text{H}_6$  =  $(2 \times 12) + (6 \times 1)$   
= 30 g

$$1 \text{ molecule of ethane} = \frac{30 \times 1}{6 \times 10^{23}} = 5 \times 10^{-23} \text{ g}$$

$$\begin{aligned} \therefore 5 \text{ molecules of ethane} &= 25 \times 10^{-23} \\ &= 2.5 \times 10^{-22} \text{ g} \end{aligned}$$

### Trial 3

1. Using the Avogadro's constant equal to  $L$ , find the number of atoms in the following masses of calcium:

(i) 200 g [5L]

(ii) 1 kg [0.0125L]

(iii) 4 g [0.1L]

### Example 13

Calculate the mole of each element present in 18.4 g of iron (III) sulphate.

### Solution

Molar mass of  $\text{Fe}_2(\text{SO}_4)_3$

$$= 2 \times 56 + 3 \times 32 + 12 \times 16$$

$$= 112 + 96 + 192$$

$$= 400 \text{ g}$$

$$\begin{aligned} \text{No. of moles of } \text{Fe}_2(\text{SO}_4)_3 &= \frac{\text{Mass given}}{\text{Mass of 1 mole}} \\ &= \frac{18.4}{400} = \mathbf{0.046} \text{ mole} \end{aligned}$$

$$\begin{aligned} \therefore \text{No. of moles of iron} &= 0.046 \times 2 \\ &= 0.092 \text{ moles.} \end{aligned}$$

$$\begin{aligned} \therefore \text{No. of moles of sulphur} &= 0.046 \times 3 \\ &= 0.138 \text{ moles} \end{aligned}$$

$$\begin{aligned} \therefore \text{No. of moles of oxygen} &= 0.046 \times 12 \\ &= 0.552 \text{ moles} \end{aligned}$$

2. Taking Avogadro's constant equal to  $6 \times 10^{23}$ , find the number of molecules in

(a) 3.4 g of ammonia,  $\text{NH}_3$  [ $1.2 \times 6 \times 10^{23}$ ]

(b) 7.7 g of carbon tetrachloride  $\text{CCl}_4$   
[ $0.05 \times 6 \times 10^{23}$ ]

(c) 16 g of sulphur dioxide,  $\text{SO}_2$   
[ $0.25 \times 6 \times 10^{23}$ ]

(d) 1 g of hydrogen, H<sub>2</sub>.

$$[0.5 \times 6 \times 10^{23}]$$

3. What mass of carbon dioxide has the same number of molecules as 7 g of nitrogen, N<sub>2</sub>? [11]
4. What is the mass of 0.5 mole of aluminium sulphate, Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub>? [171]

### Percentage composition

#### Example 14

Calculate the percentage composition by mass of magnesium oxide, (Mg = 24, O = 16).

- (i) Write formula of the compound: MgO
- (ii) Find the formula mass of the compound:  
24 + 16 = 40
- (iii) Express each atomic mass as a percentage of the formula mass.

Percentage of Magnesium

$$= \frac{24 \times 100\%}{40} = 60\%$$

Percentage of oxygen

$$= \frac{16 \times 100\%}{40} = 40\%$$

#### Example 15

Calculate the percentage of water of crystallization in hydrated sodium carbonate, Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O?

Formula mass of Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O

$$\begin{aligned} &= 46 + 12 + 48 + 20 + 160 \\ &= 286 \end{aligned}$$

**Note:** The percentage of water should be calculated as a complete, separate unit i.e. (10 H<sub>2</sub>O = 180) in this case.

per cent water in Na<sub>2</sub>CO<sub>3</sub>·10H<sub>2</sub>O

$$\begin{aligned} &= \frac{180 \times 100\%}{286} \\ &= 62.94\% \end{aligned}$$

### Empirical and molecular Formulae

The **empirical formula (E.F)** of a compound is the simplest formula which expresses its composition by mass. It does not give the exact number of various atoms present in one formula unit of molecule of a compound.

The **molecular formula (M.F)** of a compound is the one which shows the exact number of each kind of each atom present in a molecule of a compound.

The molecular formula is always a simple multiple of the empirical formula.

i.e. M.F = n(E.F) where n is a whole number e.g., E.F of glucose is CH<sub>2</sub>O, whereas its M.F. is C<sub>6</sub>H<sub>12</sub>O<sub>6</sub> which (CH<sub>2</sub>O)<sub>6</sub>.

OR  $\frac{\text{Molecular formulamass}}{\text{Emperical formulamass}} = n$ , a whole

number.

### Example 16

Determine the formula of a lead compound, given that 4.14 g of lead combines with 0.64 g of sulphur and 1.28 g of oxygen.

We shall use the 'mole' concept to find the formula of the compound. Proceed as shown in the table below:

Element symbol	Lead, Pb	Sulphur, S	Oxygen, O
Mass in g	4.14	0.64	1.28
RAM	207	32	16
No. of moles	$\frac{4.14}{207} = 0.02$	$\frac{0.64}{32} = 0.02$	$\frac{1.28}{16} = 0.08$
Ratio of mole	$\frac{0.02}{0.02} = 1$	$\frac{0.02}{0.02} = 1$	$\frac{0.08}{0.02} = 4$
No. of atoms	1	1	4
E.F	PbSO <sub>4</sub>		

### Example 17

The combustion of 0.92 g of copper gave 1.15 g of copper oxide.

Calculate the percentage of copper and oxygen in the sample and hence its formula.

Mass of oxygen in the oxide = 1.15 - 0.92

$$= 0.23 \text{ g}$$

Percentage of copper

$$= \frac{0.92 \times 100\%}{1.15} = 80\%$$

Percentage of oxygen

$$= \frac{0.23 \times 100\%}{1.15} = 20\%$$

	Copper, Cu	Oxygen, O
RAM	64	16
Percentage	80	20
No. of moles	$\frac{80}{64} = 1.25$	$\frac{20}{16} = 1.25$
No. of atoms	$\frac{1.25}{1.25} = 1$	$\frac{1.25}{1.25} = 1$
Formula	CuO	

∴ The formula of copper oxide is CuO.

### Example 18

3.22 g of hydrated sodium sulphate, Na<sub>2</sub>SO<sub>4</sub>.nH<sub>2</sub>O were heated till all the water of

crystallization was driven off. The anhydrous salt left had a mass of 1.42 g. Determine the formula of the hydrated salt.

Compound	Na <sub>2</sub> SO <sub>4</sub>	H <sub>2</sub> O
Mass in g	1.42	1.80 (3.22 - 1.42)
R.M.M	142	180
No. of moles	$\frac{1.42}{142} = 0.01$	$\frac{1.80}{18} = 0.1$
Ratio of mole	1	10

Hydrated sodium sulphate is Na<sub>2</sub>SO<sub>4</sub>.10H<sub>2</sub>O

### Example 19

An organic compound was found to contain 12.8% carbon, 2.1% Hydrogen and 85.1% bromine. A rough estimate gave its relative molecular mass between 150 and 200.

### Trial 4

1. What is the percentage of hydrogen and oxygen in ethanoic acid, CH<sub>3</sub>COOH? [53.3%]
2. 24.0 g of a lead oxide was completely reduced by heating it with carbon to 20.8 g of lead. What is the formula of the oxide that was reduced? [PbO<sub>2</sub>]
3. If 0.5 mole of a hydrated salt X.nH<sub>2</sub>O contains 63 g of water of crystallization, find the value of n.[7]

Determine its molecular formula. Explain your answer.

Element	C (=12)	H (=1)	Br (=80)
No. of mole	$\frac{12.8}{12} = 1.06$	$\frac{2.1}{1} = 2.1$	$\frac{85.1}{80} = 1.06$
/mole ratio/No. of atoms	1	2	1

Empirical formula = CH<sub>2</sub>Br

E.F. mass = 12 + 2 + 80 = 94

Molecular formula mass = n x E.F. Mass  
= n x 94

i.e. M.F. mass can be 94 or 188 or 289.

In this case, it is then 188 and the M.F. of the compound is C<sub>2</sub>H<sub>4</sub>Br<sub>2</sub>

4. 2.8 g of iron reacted with 5.3 g of chlorine to form the chloride.
  - (a) Determine the simplest formula of the chloride. [FeCl<sub>3</sub>]
  - (b) If the molecular mass of the chloride is 325. Find the molecular formula and structural formula of the chloride. [Fe<sub>2</sub>Cl<sub>6</sub>]
5. When 3.6 g of magnesium was burnt in nitrogen, 5.0 g of magnesium nitride was

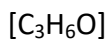
formed. Determine the simplest formula of the nitride.  $[Mg_3Cl_2]$

6. A liquid compound X has relative molecular mass of 114 and the ratio of carbon atoms to hydrogen atoms is 4:9. Determine the molecular formula of the compound. To which homologous series of organic compounds does it belong? [alkane]
7. An organic compound contains by mass 60% carbon, 13.3% hydrogen and 27.7% oxygen. Determine its simplest formula? Find its possible molecular formula if its relative molecular mass is estimated to be between 50 and 80.  $[C_3H_8O]$
8. When 0.29 g of a substance X was burnt in oxygen, 0.66 g of carbon dioxide and 0.27 g of water were formed.

Calculate

(i) the masses of carbon, hydrogen and oxygen in X.

(ii) Determine the empirical formula of X.



- 9 (a) A compound X contains carbon, 66.7% and hydrogen 11.1% the rest being oxygen. Determine the empirical formula of X.  $[C_4H_8O]$
- (b) The vapor density of X is 36; determine the molecular formula of X.  $[C_4H_8O]$
10. Combustion of 5.4 g of a compound containing C, H and O only gave 7.92 g of  $CO_2$  and 3.24 g of  $H_2O$ . Calculate the empirical formula of the compound? The

compound has a R.M.M of 180, determine its molecular formula.  $[CH_2O]$

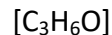
11. (a) A compound b contain 92.31% carbon and 6.69% hydrogen

Determine the empirical formula of B (2marks)  $[C_6H_6]$

(b) B burns with a sooty flame and has a vapour density of 39. Determine molecular formula and write its molecular structure.  $[C_6H_6]$

12. A compound X, vapor density 58, contain carbon 62.07%, hydrogen 10.34%, the rest being oxygen. X does not burn with sooty flame.

(i) Calculate the empirical formula of X (C=12, H=1, O=16) (3marks)



Determine the molecular formula of X  $[C_3H_6O]$

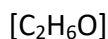
13. When 4.90g of organic compound X, containing carbon and hydrogen only was burnt in oxygen, 15.78g of carbon dioxide and 5.38g of water were formed. Calculate the empirical formula of X (2 ½ marks)  $[CH_2]$

X was steam distilled at  $80^\circ C$  and 760mmHg and the distillate was found to contain 90.8% by mass of X. (the vapor pressure of water at  $80^\circ C$  is 240mmHg).

- a. Calculate the formula mass of X. [82]  
b. Deduce the molecular formula of X.  $[C_6H_{12}]$

14. A compound P contains 52.2% carbon, 13.0% hydrogen, the rest being oxygen

(a) Determine the empirical formula of P.



(b) When vaporized, 0.1g of P occupied  $78.8\text{cm}^3$  at  $157^\circ\text{C}$  and pressure of 740mmHg.

- (i) Calculate the formula mass of P
- (ii) Determine the molecular formula of P [46]
- (iii) Write the structural formula of all the possible isomers of P
- (iv) P does not react with sodium metal. Identify P
- (v) Write equation to show how P can be prepared from methanol.

### Calculation involving Equations, Masses and Volumes

Chemical changes can be described by use of chemical equations, e.g., carbon dioxide reacts with oxygen to form carbon dioxide. This reaction is written as:



The plus sign (+) indicates addition of carbon to oxygen (vice versa) OR a mixture of the two. The arrow ( $\rightarrow$ ) is often read 'yields'. Substances on the left of the arrow are **reactants, or starting materials**. Those on the right are the **products** of the reaction.

Chemical equations have meaning on the atomic and molecular level. The equation (1) means that one atom of carbon (C) reacts with one molecule of oxygen ( $O_2$ ) to produce one molecule of carbon dioxide,  $CO_2$ . One carbon atom weighs 12 grams. One oxygen molecule weighs 32 grams, (2 x 16 grams per oxygen atom).

Similarly, in the above equation 12 grams of carbon react with 32 grams of oxygen. i.e., 12 mass units of carbon react with 32 mass units

of oxygen. The ratio of the masses of the individual reactants is 12:32.

It is not convenient to weigh gases. They are normally measure by volumes. The volume occupied at s.t.p. by one mole of oxygen or carbon dioxide is  $22.4\text{ dm}^3$  or  $22400\text{ cm}^3$ . This volume is called the molar volume.

*Equal volumes of all gases, under the same conditions of temperature and pressure, contain the same number of molecules (Avogadro's Law)*

*It follows that 1 mole of any gas occupies a volume of  $22.4\text{ dm}^3$  at s.t.p. and a volume of  $24.0\text{ dm}^3$  at r.t.p. (room temperature and pressure).*

These are the values we will be using in all calculations.

### Gay-Lussac's Law

*When gases react together, they do so in volumes which are related to each other in*

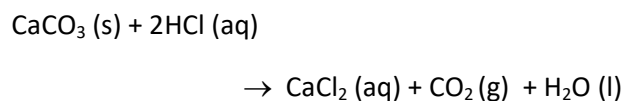
simple ratio and to the product if gaseous, provided all measurements are made at the same temperature and pressure.

### Example 20

How much marble ( $\text{CaCO}_3$ ) remains undissolved and what volume of carbon dioxide (at r.t.p.) formed when all 14.6 g of hydrochloric acid are allowed to react with 50g marble?

#### Solution

The equation is as follows:



1 mole      2 moles                      1 mole

The question can be solved using mole concept.

Molar mass mole of HCl = 36.5 g.

$$\text{Moles of HCl used} = \frac{14.6}{36.5} = 0.4 \text{ mole}$$

Mass of  $\text{CaCO}_3$  that reacted (thus dissolved)

From the equation,

2 moles of HCl react with 1 mole of  $\text{CaCO}_3$  (100)

$$\begin{aligned} 0.4 \text{ mole of HCl react with } & \frac{0.4}{2} \\ & = 0.2 \text{ mole of } \text{CaCO}_3. \\ & = (0.2 \times 100)\text{g} \\ & = 20 \text{ g of } \text{CaCO}_3 \end{aligned}$$

$$\begin{aligned} \therefore \text{Mass of undissolved } \text{CaCO}_3 & = (50 - 20) \\ & = 30 \text{ g.} \end{aligned}$$

Volume of  $\text{CO}_2$

2 moles of HCl liberate 1 mole or 24 dm<sup>3</sup> of  $\text{CO}_2$  at r.t.p.

$$\begin{aligned} \therefore 0.4 \text{ mole of HCl will liberate } & \frac{0.4 \times 24}{2} \\ & = 4.8 \text{ dm}^3 \text{ of } \text{CO}_2 \end{aligned}$$

### Example 21

According to the equation:  $2\text{C} + \text{O}_2 \rightarrow 2\text{CO}$

How many moles of carbon which react with 9.6 dm<sup>3</sup> of oxygen at r.t.p?

#### Solution

1 mole or 24 dm<sup>3</sup> (at r.t.p.) of  $\text{O}_2$  react with 2 mole of carbon.

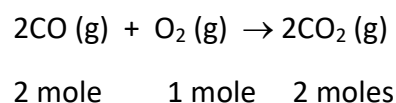
1 dm<sup>3</sup> of  $\text{O}_2$  reacts with  $\frac{2}{24}$  moles of carbon

$\therefore$  9.6 dm<sup>3</sup> react with  $\frac{2 \times 9.6}{24} = 0.8 \text{ moles}$  of carbon.

### Example 22

40 cm<sup>3</sup> of oxygen was added to 30 cm<sup>3</sup> of carbon monoxide and the mixture ignited. What is the volume and composition of the resulting mixture? (All volumes measured at the same temperature and pressure).

#### Solution



2 vol.      1 vol.      2 vol. (Gay-Lussac's Law)  
 30 cm<sup>3</sup>      15 cm<sup>3</sup>      30 cm<sup>3</sup> (volume ratio)

The volume of CO<sub>2</sub> produced is the same as that of CO used up, i.e. 30 cm<sup>3</sup> of CO<sub>2</sub> are produced.

15 cm<sup>3</sup> of the 40 cm<sup>3</sup> of oxygen are used up in the reaction, i.e. 25 cm<sup>3</sup> of oxygen remained

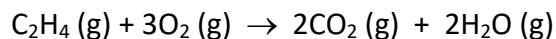
30 cm<sup>3</sup> of carbon dioxide produced and thus has a volume of 55 cm<sup>3</sup>.

### Trial 5

- Calculate the mass of magnesium oxide obtained by burning 2.4g of magnesium in excess oxygen. [4g]
- 0.2 g of copper (II) oxide is heated in a stream of hydrogen. Calculate the mass of copper obtained if the reduction is 100 percent. [0.16g]
- What mass and volume of oxygen (at r.t.p.) would be formed when 4.9 g of potassium chlorate (KClO<sub>3</sub>) are heated and completely decomposed into potassium chloride and oxygen? [1.44dm<sup>3</sup>]
- Find the mass of sodium carbonate produced and the volume of carbon dioxide gas evolved at r.t.p. when 42 g of sodium hydrogen carbonate (NaHCO<sub>3</sub>) is heated and completely decomposed? [26.5g, 6dm<sup>3</sup>]
- Write the equation of the reaction between sodium and water. 4.6 g of sodium are completely dissolved in water. Calculate the mass and volume (at r.t.p.) of hydrogen produced. [0.2g, .4dm<sup>3</sup>]
- When 3.1 g of a carbon dioxide MCO<sub>3</sub> are heated to constant mass, the volume of carbon dioxide evolved during the heating is 600 cm<sup>3</sup> at r.t.p. What is the relative atomic mass of M? [64]
- 2.8 g of iron is reacted completely with chloride according to the reaction:  
 $2 \text{Fe (s)} + 3 \text{Cl}_2 \text{(g)} \rightarrow 2 \text{FeCl}_3 \text{(s)}$   
 Find the mass of iron (III) chloride formed and the minimum volume of chloride at r.t.p. needed for this reaction. [8.13g, 1.8dm<sup>3</sup>]
- Find the mass of copper (II) sulphide that can be precipitated when hydrogen sulphide is bubbled through a solution of copper (II) sulphate containing 4.0 g of copper (II) sulphate. [2.35g]
- The equation for the reduction process taking in the blast furnace is  
 $\text{Fe}_2\text{O}_3\text{(s)} + 3\text{CO (g)} \rightarrow 2\text{Fe(s)} + 3\text{CO}_2 \text{(g)}$   
 What is the maximum mass of iron which could be extracted from 40 tons of iron (III) oxide? What is the mass and volume (at r.t.p) of carbon dioxide produced? [27.97 tons, 2.41 x 10<sup>8</sup> dm<sup>3</sup>]
- What mass of 90 percent pure ammonium chloride would be needed just to react completely with 14.8 g of slaked lime, Ca(OH)<sub>2</sub>? What volume at s.t.p of ammonia will be produced? [23.73g]
- What volume of oxygen is need for complete combustion of 20 litres of

propane (C<sub>3</sub>H<sub>8</sub>) to carbon dioxide and water, measured at r.t.p.? [100L]

12. Ethene reacts with oxygen according to the equation



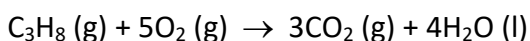
15 cm<sup>3</sup> of ethene were mixed with 60 cm<sup>3</sup> of oxygen and the mixture was sparked to complete the reaction. All volumes were measured at a pressure of 1 atm. and (i) at a temperature of 120<sup>o</sup> C (ii) at room temperature.

Find the volume and composition of the gas mixture left at the end of the experiment in each case.

[at room temperature → 45dm<sup>3</sup>, at 1200 → 75dm<sup>3</sup>]

13. 10 cm<sup>3</sup> of a gaseous hydrocarbon were mixed with 90 cm<sup>3</sup> of oxygen and sparked. The resulting volume at r.t.p. was 70 cm<sup>3</sup>, which reduced to 30 cm<sup>3</sup> on shaking with sodium hydroxide. Find the empirical formula of the hydrocarbon. [C<sub>4</sub>H<sub>8</sub>]

14. According to the equation,



- (a) How many moles of CO<sub>2</sub> will be produced in the reaction of 1 Mole of O<sub>2</sub>? [0.6moles]  
(b) How many grams of CO<sub>2</sub> are produced by 1 mole O<sub>2</sub>? [26.4g]  
(c) How many grams of propane (C<sub>3</sub>H<sub>8</sub>) will react with 3.2 g of O<sub>2</sub>? [0.88g]  
(d) How many moles of propane (C<sub>3</sub>H<sub>8</sub>) will produce 9 g of H<sub>2</sub>O? [0.125moles]

- (e) How many moles of CO<sub>2</sub> will be produced with 36 g of water? [1.5moles]

- (f) What are the maximum moles of CO<sub>2</sub> that will be produced from a mixture of 12 g of C<sub>3</sub>H<sub>8</sub> and 8 g of O<sub>2</sub>? Assume complete reaction. [0.15moles]

15. When 20 g of a mixture containing anhydrous sodium carbonate and sodium hydrogen carbonate was heated and cooled, the mass of the mixtures changed to 13.8 g.

- (a) Write equation for the reaction that took place during the heating (1½ marks)

- (b) Calculate the percentage of sodium carbonate in the mixture [16%]

16. 1.18g of compound P on evaporation occupied 300cm<sup>3</sup> at s.t.p.

- (a) Calculate the relative molecular mass of P. [88]

- (b) The empirical formula of P is C<sub>2</sub>H<sub>4</sub>O. Determine its molecular formula [C<sub>4</sub>H<sub>8</sub>O<sub>2</sub>]

- (c) Compound P reacts with sodium carbonate to produce a gas that turns lime water milky. Write the structural formula of P. [CH<sub>3</sub>CH<sub>2</sub>CH<sub>2</sub>COOH]

17. When 0.55g of organic compound T was burned in oxygen 0.781g of carbon dioxide and 0.479g of water were formed.

- (i) Determine empirical formula of T [CH<sub>3</sub>O]

- (ii) When 0.225g of T was vaporized at  $127^{\circ}\text{C}$  and 760mmHG it occupied a volume of  $119.11\text{cm}^3$ . (Molar volume of a gas at s.t.p is  $22.4\text{dm}^3$ ). calculate the formula mass of T and determine molecular formula of T [ $62, \text{C}_2\text{H}_6\text{O}_2$ ]
- (iii) T reacts with acidified potassium dichromate to form ethane-1,2-dioic acid. Write the formula and the IUPAC name of T. [ethane-1,2-diol]
18. When  $142\text{cm}^3$  of a hydrocarbon Q, of molecular mass 58 was exploded with excess oxygen and cooled to room temperature, the volume of the residual gas was  $694\text{cm}^3$ . After addition of concentrated potassium hydroxide, the volume decreased to  $126\text{cm}^3$
- (a) determine the molecular formula of Q [ $\text{C}_4\text{H}_{10}$ ]
- (b) write the names and the structural formula of all isomers of Q
19. (a) When 8.8g of a hydrocarbon, Z, was burn in excess air, 14.4g of water and  $13.44\text{dm}^3$  of carbon dioxide were obtained at s.t.p. determine the empirical formula of Z.
- (a) The vapor density of Z is 22. Write the name and molecular formula of Z. [ $\text{C}_3\text{H}_8$ , propane]
- (b) Write equations to show how Z can be synthesized from alcohol
- (c) Indicate a mechanism for the first stage of the reaction in (c)(i).
20. A compound P contains 52.2% carbon, 13.0% hydrogen, the rest being oxygen
- (a) Determine the empirical formula of P. [ $\text{C}_2\text{H}_6\text{O}$ ]
- (b) When vaporized, 0.1g of P occupied  $78.8\text{cm}^3$  at  $157^{\circ}\text{C}$  and pressure of 740mmHg.
- (i) Calculate the formula mass of P [46g]
- (ii) Determine the molecular formula of P [ $\text{C}_2\text{H}_5\text{OH}$ ]
- (iii) Write the structural formula of all the possible isomers of P
- (iv) P does not react with sodium metal. Identify P
- (v) Write equation to show how P can be prepared from methanol.

### Calculations involving Concentration of Solution.

The **concentration** of a substance is the number of moles or the mass of a solute dissolved or contained in a known volume of solution. Usually the concentration of a solution is expressed in either number of grams or number of moles of solute per litre of solution.

A **standard substance/primary standard** is a pure substance whose mass can easily be weighed accurately and is used to standardize others.

### Properties of primary standard

- Should have high molecular mass to be weighed accurately with little errors

- Should form stable solutions
- It should be stable in air, i.e., it shouldn't be hygroscopic, deliquescent, react with air, easily lose its water of crystallization and/or it should not be volatile.

A **Standard solution** is a solution whose concentration is known, usually expressed in terms of molarity, mole/litre, moles/dm<sup>3</sup>, or grams/litre.

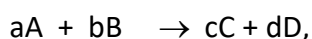
The **Molarity** of solution is the number of moles of the solute contained in 1 litre, 1000 ml or 1000 cm<sup>3</sup> of solution.

$$\text{Molarity} = \frac{\text{mass of solute dissolved in one litre}}{\text{molar (formula) mass}}$$

$$\text{or} = \frac{\text{concentration in g/l}}{\text{molar (formula) mass}}$$

$$\text{Hence, No. of moles} = \frac{\text{Molarity} \times \text{volume}}{1000 \text{ ml of solution}}$$

If solutions of compounds A and B react according to the equation



$$\text{then } \frac{V_A \times M_A}{V_B \times M_B} = \frac{a}{b}$$

where, V<sub>A</sub> and V<sub>B</sub>: volumes of solution

and m<sub>A</sub> and m<sub>B</sub>: molarities of solutions

A **molar solution** is a solution containing 1 mole of a solute/substance per litre.

### Example 23

Calculate the molarities of

(a) 7g of potassium hydroxide KOH, in 250cm<sup>3</sup> solution

$$\text{Mass of KOH in g/l} = \frac{1000 \times 7}{250} = 28\text{g/l}$$

$$\text{RFM of KOH} = (39.1 + 16 + 1) = 56.1$$

$$\text{Mole of KOH per litre} = \frac{28}{56.1} = 0.5$$

$$\text{Molarity} = 0.5\text{M}$$

(b) 9.8g of sulphuric acid in 2 dm<sup>3</sup> solution.

$$\text{RMM of H}_2\text{SO}_4 = 2 + 32 + 64 = 98\text{g}$$

$$\begin{aligned} 9.8\text{g of H}_2\text{SO}_4 \text{ in } 2\text{dm}^3 \text{ solution} \\ = 4.9\text{g in } 1\text{ dm}^3 \text{ solution.} \end{aligned}$$

$$\text{Mole of H}_2\text{SO}_4 \text{ per litre} = \frac{4.9}{98} = 0.05$$

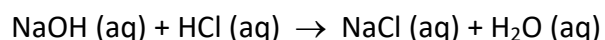
$$\therefore \text{Molarity of H}_2\text{SO}_4 = 0.05\text{M}$$

### Example 24

25.0 cm<sup>3</sup> of a solution of sodium hydroxide of concentration 0.1 mole dm<sup>-3</sup> were exactly neutralized by 20.0 cm<sup>3</sup> of a solution of hydrochloric acid. Calculate the concentration of the acid solution (a) as a molarity (b) in g/dm<sup>3</sup>.

### Solution

Reaction equation:



Mole of NaOH

1000 cm<sup>3</sup> contain 0.1 moles

25 cm<sup>3</sup> will contain  $\frac{25 \times 0.1}{1000} = 0.0025 \text{ moles}$

Moles of HCl

From the equation

1mole of NaOH react with 1 mole of HCl

⇒ moles of HCl = moles of NaOH  
= 0.0025moles

Molarity of HCl

20cm<sup>3</sup> contain 0.0025mole

1000cm<sup>3</sup> will contains  $\frac{0.0025 \times 1000}{20}$   
=0.125moles

=> molarity of HCl = 0.125M

Concentration = moles x RFM

= 0.125 x 36.5

= 4.58 g/dm<sup>3</sup>

### Example 25

In a titration 30 cm<sup>3</sup> of 0.4M sodium hydroxide require 40cm<sup>3</sup> of 0.15 M phosphoric acid, H<sub>3</sub>PO<sub>4</sub>.

(i) How many moles of sodium hydroxide are present in 30cm<sup>3</sup> of solution?

1000cm<sup>3</sup> contain 0.4 moles

30cm<sup>3</sup> contain  $\frac{30 \times 0.4}{1000} = 0.012 \text{ moles}$

(ii) How many moles of phosphoric acid are present in 40cm<sup>3</sup> of solution?

1000cm<sup>3</sup> contain 0.15 moles

40cm<sup>3</sup> contain  $\frac{40 \times 0.15}{1000} = 0.006 \text{ moles}$

(iii) Find the number of moles of sodium hydroxide required to neutralize one mole of phosphoric acid?

0.006mole of H<sub>3</sub>PO<sub>4</sub> → 0.012mole of NaOH

∴ 1mole of H<sub>3</sub>PO<sub>4</sub> →  $\frac{0.012}{0.006}$

= 2 mole of NaOH

(iv) Write the formula of the sodium salt formed and the equation for the reaction between the acid and sodium hydroxide?

2NaOH (aq) + H<sub>3</sub>PO<sub>4</sub> (aq) →

Na<sub>2</sub>HPO<sub>4</sub>(aq) + 2H<sub>2</sub>O(l)

### Example 26

24.0cm<sup>3</sup> of sodium of 0.1M KOH were exactly neutralized by 30.0cm<sup>3</sup> of sodium of χ M H<sub>2</sub>SO<sub>4</sub>. Find the concentration of the acid solution (a) as the molarity χ (b) in g/dm<sup>3</sup>

2KOH(aq) + H<sub>2</sub>SO<sub>4</sub>(aq) → K<sub>2</sub>SO<sub>4</sub> (aq) + 2H<sub>2</sub>O(l)

Mole of KOH

1000 cm<sup>3</sup> contain 0.1 moles

24.0 cm<sup>3</sup> will contain  $\frac{24 \times 0.1}{1000} = 0.0024 \text{ moles}$

Moles of  $\text{H}_2\text{SO}_4$

From the equation

2mole of KOH react with 1 mole of  $\text{H}_2\text{SO}_4$

$$\begin{aligned} \Rightarrow \text{moles of } \text{H}_2\text{SO}_4 &= \frac{1}{2} \text{ moles of KOH} \\ &= \frac{0.0024}{2} = 0.0012 \text{ moles} \end{aligned}$$

Molarity of  $\chi$  of  $\text{H}_2\text{SO}_4$

$30\text{cm}^3$  contain  $\overline{0.0012\text{mole}}$

$1000\text{cm}^3$  will contains

$$\frac{0.0012 \times 1000}{30} = 0.04 \text{ moles}$$

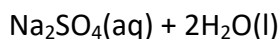
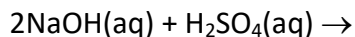
$$\therefore \chi = 0.04M$$

Thus  $\text{g/dm}^3$  of  $\text{H}_2\text{SO}_4 = 0.04 \times 98 = 3.92\text{g/dm}^3$

### Example 27

*In an experiment to determine the concentration dilute sulphuric acid in moles per litre, 25 cm<sup>3</sup> of 0.2M sodium hydroxide solution required 24.6 of the acid. Calculate the molarity of the acid.*

Equation.



Moles of the NaOH that reacted

$1000\text{cm}^3$  contains 0.2 mole

$1\text{cm}^3$  contains  $\frac{0.2}{1000}$  moles

$25\text{cm}^3$  contain  $\frac{0.2 \times 25}{1000} = 0.005 \text{ moles}$

Moles of  $\text{H}_2\text{SO}_4$

From the equation: 2 mole of NaOH reacts with 1 mole  $\text{H}_2\text{SO}_4$ .

0.05 mole of NaOH react with  $(\frac{1}{2} \times 0.005)$  mole of acid

Molarity

$$\begin{aligned} \therefore \text{Mole of } \text{H}_2\text{SO}_4 \text{ in } 24.6\text{cm}^3 \text{ that reacted} \\ = 0.0025 \text{ mole} \end{aligned}$$

$1\text{cm}^3$  contain  $\frac{0.0025}{24.6}$

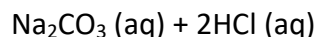
$$\begin{aligned} \Rightarrow 1000\text{cm}^3 \text{ contain} \\ \frac{0.0025 \times 1000}{24.6} = 0.102 \text{ mole} \end{aligned}$$

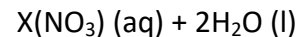
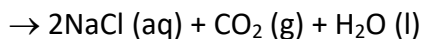
Therefore, molarity of  $\text{H}_2\text{SO}_4$  solution is **0.102M**

### Example 28

*Find the volume of 0.25M HCl required to exactly react with 20 cm<sup>3</sup> of 0.1M sodium carbonate solution.*

Reaction equation





Moles of the  $\text{Na}_2\text{CO}_3$  that reacted

$1000 \text{ cm}^3$  contains 0.1 moles

$1 \text{ cm}^3$  contains  $\frac{0.1}{1000}$  moles

$20 \text{ cm}^3$  contain  $\frac{0.1 \times 20}{1000} = 0.002 \text{ moles}$

Moles of HCl \_\_\_\_\_

From the equation; 1 mole of  $\text{Na}_2\text{CO}_3$  reacts with 2 mole HCl

0.02 mole of  $\text{Na}_2\text{CO}_3$  react with  $(0.002 \times 2)$  mole of HCl = 0.004 moles

Volume of hydrochloric acid required

0.25 moles are contained in  $1000 \text{ cm}^3$

$\Rightarrow 0.004$  moles will be in  $\frac{1000 \times 0.004}{0.25}$   
=  $16 \text{ cm}^3$

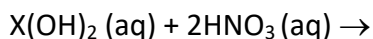
The volume required is  **$16 \text{ cm}^3$**

### Example 29

$20 \text{ cm}^3$  of  $0.1 \text{ M}$  nitric acid was completely neutralized by  $10 \text{ cm}^3$  of a hydroxide of a divalent metal  $\text{X(OH)}_2$  which was made by dissolving  $5.8 \text{ g}$  of the hydroxide in 1 litre of solution. ( $H=1, O=16$ ).

(a) Calculate the molarity of the hydroxide.

Reaction equation



No. of moles of the nitric acid that reacted

$1000 \text{ cm}^3$  contains 0.1 moles

$1 \text{ cm}^3$  contains  $\frac{0.1}{1000}$  moles

$\Rightarrow 20 \text{ cm}^3$  contain  $\frac{20 \times 0.1}{1000}$   
= 0.002 moles

(b) Calculate the molar mass of  $\text{X(OH)}_2$  and hence determine the atomic mass of X.

From the reaction equation 2 moles of nitric acid react with 1 mole of the hydroxide.

1 mole of nitric acid react with  $\frac{1}{2}$  moles of  $\text{X(OH)}_2$

$\Rightarrow$  moles of the hydroxide in reacted  $0.002 \times \frac{1}{2}$   
= 0.001 mole

Molarity of  $\text{X(OH)}_2$

$\therefore 10 \text{ cm}^3$  of  $\text{X(OH)}_2$  contain 0.001 mole

$1 \text{ cm}^3$  contains  $\frac{0.001}{1000}$  moles

$\Rightarrow 1000 \text{ cm}^3$  contains  $\frac{0.001 \times 1000}{10}$   
= 0.1 moles

Therefore, molarity of the hydroxide = **0.1M**.

(b) Molar mass of  $X(OH)_2 =$

$$\frac{\text{concentration g/l}}{\text{molarity}}$$
$$= \frac{5.8}{0.1} = 58g$$

The atomic mass of X is obtained from using the molar mass, i.e.

$$\text{Molar mass of } X(OH)_2 = X + 2(16+1) = 58$$

$$= X + 34 = 58$$

$$\Rightarrow X = 24g$$

Thus, the hydroxide of compound X is magnesium hydroxide  $Mg(OH)_2$ .

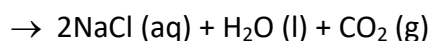
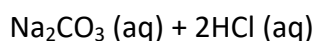
### Example 30.

In determination of water of crystallization of  $Na_2CO_3 \cdot nH_2O$ ,  $14 \text{ cm}^3$  of  $0.2M$  HCl was required to neutralize  $25 \text{ cm}^3$  of sodium carbonate prepared by dissolving  $4g$  of  $Na_2CO_3 \cdot nH_2O$  in  $250 \text{ cm}^3$  of solution.

(a) Calculate the molarity of sodium carbonate.

(b) Calculate the molar mass of the hydrated sodium carbonate and hence calculate the number of moles n of water of crystallization.

Equations.



Moles of the acid that reacted

$$1000 \text{ cm}^3 \text{ contain } 0.2 \text{ mole}$$

$$1 \text{ cm}^3 \text{ contain } \frac{0.2}{1000} \text{ moles}$$

$$\therefore 14 \text{ cm}^3 \text{ contain } \frac{0.2 \times 14}{1000}$$

$$= 0.0028 \text{ moles}$$

From the reaction equation, 2 moles of HCl react with 1 mole of the base.

Thus, moles  $Na_2CO_3$  in  $25 \text{ cm}^3$  that reacted

$$= \frac{0.0028}{2}$$

$$= 0.0014 \text{ moles}$$

$$1 \text{ cm}^3 \text{ contain } = \frac{0.28}{2 \times 25}$$

$$\Rightarrow 1000 \text{ cm}^3 \text{ contain } \frac{0.28 \times 1000}{2 \times 25}$$

$$= 0.056 \text{ M.}$$

thus, molarity of sodium carbonate solution is  $0.056M$ .

(b) Molar mass of hydrated sodium carbonate

$$\text{Molar mass} = \frac{\text{Concentration g/l}}{\text{Molarity}}$$

$$= \frac{2 \times 4}{0.056}$$

$$= 286 \text{ g}$$

Value of n in the formula  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$  is calculated using the molar mass.

$$\text{thus, } (23 \times 2) + 12 + (16 \times 3) + 18n = 286$$

$$18n = 286 - 106$$

$$n = 10$$

### Example 31

During determination of Basicity of an acid  $\text{H}_n\text{A}$ ,  $25 \text{ cm}^3$  of KOH prepared by dissolving 11.2 g in 1 litre of solution was required to neutralize  $25 \text{ cm}^3$  of 0.1M  $\text{H}_n\text{A}$ .

(a) Calculate the molarity of the hydroxide.

$$\begin{aligned} \text{Formula mass of KOH} &= \frac{39.1 + 16 + 1}{1} \\ &= 56.1 \end{aligned}$$

$$26.1 \text{ g of KOH} \rightarrow 1 \text{ mole}$$

$$\therefore 11.2 \text{ g of KOH} \rightarrow \frac{11.2}{56.1} = 0.2 \text{ moles}$$

$$\text{Hence molarity of KOH} = 0.2 \text{ M}$$

(b) Calculate the basicity of the acid.

Mole of the KOH that reacted

$$1000 \text{ cm}^3 \text{ contain } 0.2 \text{ mole}$$

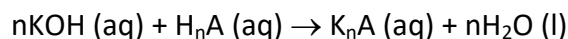
$$25 \text{ cm}^3 \text{ contain } \frac{0.2 \times 25}{1000} = 0.005 \text{ moles}$$

Mole of the acid that reacted

$$1000 \text{ cm}^3 \text{ contain } 0.1 \text{ mole}$$

$$25 \text{ cm}^3 \text{ contain } \frac{0.1 \times 25}{1000} = 0.0025 \text{ moles}$$

Equation.,



Reaction ratio n:1

$$\frac{n}{1} = \frac{0.005}{0.0025} = 2$$

Basicity, n = 2

### Example 32

What volume of 5M  $\text{H}_2\text{SO}_4$  is required to prepare  $250 \text{ cm}^3$  of 0.25M of the acid.

**Solution**

$$\text{Molarity}^1 \times \text{Volume}^1 = \text{Molarity}^2 \times \text{Volume}^2$$

$$5 \times \text{volume conc. acid} = 0.25 \times 250$$

$$\begin{aligned} \text{Volume of the conc. acid} &= \frac{0.25 \times 250}{5} \\ &= 12.5 \text{ cm}^3 \end{aligned}$$

Or

Moles of sulphuric acid in  $250 \text{ cm}^3$  of 0.25M solution

$$1000 \text{ cm}^3 \text{ contain } 0.25 \text{ moles}$$

$$\Rightarrow 250 \text{ cm}^3 \text{ contain } \frac{0.25 \times 250}{1000} = 0.0625 \text{ moles}$$

Volume of concentrated acid that contain 0.0625 moles

$$5 \text{ moles are contained in } 1000 \text{ cm}^3$$

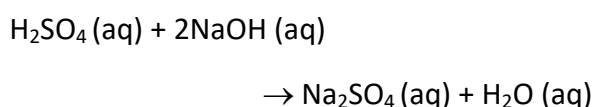
$$\Rightarrow 0.625 \text{ moles are contained in } \frac{0.0625 \times 1000}{5}$$

$$=12.5\text{cm}^3$$

### Example 33

10 cm<sup>3</sup> of concentrated sulphuric acid were diluted to 1000 cm<sup>3</sup>. 25 cm<sup>3</sup> of this diluted acid were neutralized by 30 cm<sup>3</sup> of 0.3 M sodium hydroxide. Calculate the mass of the acid in 1 dm<sup>3</sup> of the concentrated acid.

Equation.



Moles of NaOH that reacted.

1000 cm<sup>3</sup> contain 0.3 mole

$$1 \text{ cm}^3 \rightarrow \frac{0.3}{1000}$$

$$\therefore 25 \text{ cm}^3 \rightarrow \frac{0.3 \times 30}{1000} = 0.009 \text{ moles}$$

but, 2 moles of NaOH react with 1 mole of the acid.

$\therefore$  moles of H<sub>2</sub>SO<sub>4</sub> in 25 ml of the dilute acid

$$= 0.009 \times \frac{1}{2}$$

$$1000\text{cm}^3 \text{ contain } \frac{0.009 \times 1000}{25 \times 2} = 0.18 \text{ moles}$$

$\therefore$  molarity of the dilute solution is 0.18M

Since 10 cm<sup>3</sup> of the concentrated acid was diluted to 1000 cm<sup>3</sup> to make the dilute solution, then.

$$\text{Molarity}^1 \times \text{Volume}^1 = \text{Molarity}^2 \times \text{Volume}^2$$

$$10x = 1000 \times 0.18$$

where x is the molarity of the concentrated acid.

$$\text{thus, } X = 18\text{M}$$

The mass of the acid in = Molarity x  
Molar mass

$$\begin{aligned} 1\text{dm}^3 \text{ of concentrated acid} &= 18 \times 98 \\ &= 1764 \text{ g.} \end{aligned}$$

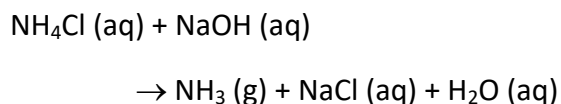
### Example 34.

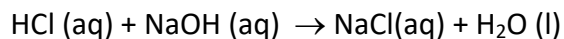
In the determination of the percentage purity of NH<sub>4</sub>Cl; 5.0 g of an impure salt of NH<sub>4</sub>Cl was boiled with 100cm<sup>3</sup> of 2M NaOH for about 10 minutes. The residual solution was made up to 250 cm<sup>3</sup> in a volumetric flask with distilled water. 25 cm<sup>3</sup> portion of the resultant solution required 22.4 cm<sup>3</sup> of 0.5M HCl for neutralization using methyl orange indicator.

(a) Calculate the percentage of NH<sub>4</sub>Cl in the sample

#### Solution

Sodium hydroxide reacts with NH<sub>4</sub>Cl to release ammonium gas which is lost on heating as per the equation.





Moles of the acid that reacted with excess base in 25cm<sup>3</sup> of the resultant solution

$$1000 \text{ cm}^3 \text{ contain } 0.5 \text{ moles}$$

$$\therefore 22.4 \text{ cm}^3 \text{ contain } \frac{0.5 \times 22.4}{1000}$$

$$= 0.0112 \text{ moles}$$

But, 1 mole of the acid reacts with 1 mole of the base

$\therefore$  moles of excess NaOH in 25 cm<sup>3</sup> of the resultant solution

$$250 \text{ cm}^3 \text{ will contain } \frac{0.0112 \times 250}{25}$$

$$= 0.0112 \text{ moles}$$

Total moles of NaOH added to impure NH<sub>4</sub>Cl

1000 cm<sup>3</sup> contain 2 moles

$$\therefore 100 \text{ cm}^3 \text{ contain } \frac{2 \times 100}{1000} = 0.2 \text{ moles}$$

Mole of NaOH that reacted with NH<sub>4</sub>Cl is equivalent to the difference between moles NaOH added and that found in excess.

$$= 0.2 - 0.112$$

$$= 0.088 \text{ moles}$$

Since 1 mole of NaOH reacts with 1 mole NH<sub>4</sub>Cl, then mole of NH<sub>4</sub>Cl that reacted

$$= 0.088$$

Mass of 0.088 moles of NH<sub>4</sub>Cl

$$= \text{molarity} \times \text{molar mass}$$

$$= 0.088 (14+4+35.5)$$

$$= 4.708\text{g}$$

Percentage purity

$$= \frac{\text{Mass of pure substance} \times 100}{\text{Total mass of impure sample}}$$

$$\therefore \text{Percentage purity of NH}_4\text{Cl} = \frac{4.708 \times 100}{5.0}$$

$$= 94.2\%$$

(b) Hence determine the percentage of ammonia in the pure sample.

Percentage of ammonia in the pure sample

Every 53.5 g of NH<sub>4</sub>Cl contain 17g of NH<sub>3</sub>

$$\therefore 4.708 \text{ contain } \frac{17 \times 4.708}{53.5} = 1.496\text{g}$$

Percentage of NH<sub>3</sub> in pure sample

$$= \frac{1.496 \times 100}{4.708}$$

$$= 31.77\%$$

## Trial 6

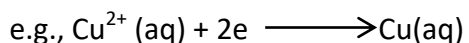
1. What is the molarity of 21.2 g of sodium carbonate dissolved in 10 litres of solution (Na=23, C=12, O=16).[0.02M]
2. Find the volume of 0.01M sodium hydroxide which is required to react exactly with 25 cm<sup>3</sup> of 0.02M HCl. [50cm<sup>3</sup>]
3. 25 cm<sup>3</sup> of 0.1M sodium carbonate was found to react with 23.3 cm<sup>3</sup> of HCl to complete neutralization. Calculate the molarity of hydrochloric acid. [0.2146M]
4. A solution of 0.1M HCl and volume 25 cm<sup>3</sup> required 12.5 cm<sup>3</sup> of a metal hydroxide M(OH)<sub>2</sub> which was prepared by dissolving 1.45 g of the M(OH)<sub>2</sub> in 250cm<sup>3</sup> of solution. (H=1, O=16)
  - (a) Calculate the molarity of M(OH)<sub>2</sub>[0.1M]
  - (b) Calculate the molar mass of M(OH)<sub>2</sub> [58]
  - (c) Calculate the atomic mass of M. [24]
5. Calculate the molarity and concentration in grams per litre of a sodium hydroxide solution, if 12.5 cm<sup>3</sup> of 0.2M H<sub>2</sub>SO<sub>4</sub> acid required 25 cm<sup>3</sup> of NaOH for complete neutralization. [0.2M, 8g/l]
6. 25 cm<sup>3</sup> of oxalic (H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>.nH<sub>2</sub>O) acid solution prepared by dissolving 2.52 g of the acid into 250 cm<sup>3</sup> required 20 cm<sup>3</sup> of sodium hydroxide solution prepared by dissolving 8 g of the hydroxide in one litre of solution.
  - (a) Calculate the molarity of sodium hydroxide solution. [0.2M]
  - (b) Calculate the molarity of oxalic (H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>.nH<sub>2</sub>O) acid solution [0.08M]
  - (c) Calculate the value of n (the mole of molecules of water of crystallization (c=12, H=1, O=16, N=23) [2]
7. Calculate the molarities/concentrations of the following solutions.
  - (a) 20.0g of sodium hydroxide, NaOH, in 250cm<sup>3</sup> of solution; [2M]
  - (b) 3.8g of iron (II) sulphate, FeSO<sub>4</sub>, in 100cm<sup>3</sup> solution.[0.25M]
8. Calculate the number of moles of the named substance in
  - (a) 50cm<sup>3</sup> of 0.1 M nitric acid. [0.005moles]
  - (b) 100cm<sup>3</sup> of 1.5 M sodium carbonate. [0.15moles]
9. Calculate the masses of dissolved substances contained in
  - (a) 100cm<sup>3</sup> of 2 M sodium sulphate solution, Na<sub>2</sub>SO<sub>4</sub>. [28.4g]
  - (b) 50cm<sup>3</sup> of 0.2 M potassium carbonate, K<sub>2</sub>CO<sub>3</sub>. [1.38g]

10. 8.48 grams of a mixture of sodium chloride and anhydrous sodium carbonate were dissolved in  $100 \text{ cm}^3$  of solution.  $5 \text{ cm}^3$  of this solution required  $40 \text{ cm}^3$  of  $0.1 \text{ M HCl}$  for complete neutralization. Calculate the percentage composition of the original mixture. [50%]
11. Concentrated hydrochloric acid containing 36% by mass of the acid has a density of  $1.18 \text{ g/cm}^3$ . Calculate the volume of this acid required to neutralize 8.4 grams of sodium hydrogen carbonate. [ $8.5 \text{ cm}^3$ ]
12.  $20 \text{ cm}^3$  of  $0.9 \text{ M}$  sodium hydroxide solution were added to  $50 \text{ cm}^3$  of nitric acid. The excess acid neutralized  $20 \text{ cm}^3$  of  $1.1 \text{ M}$  alkali. Calculate the molarity of the original acid solution and the number of grams of nitric acid per  $\text{dm}^3$  of this solution. [ $0.8 \text{ M}$ ]
13. 2 grams of impure calcium carbonate were added to  $80 \text{ cm}^3$  of  $1.25 \text{ M}$  hydrochloric acid. The resulting solution was made up to  $100 \text{ cm}^3$  and  $10 \text{ cm}^3$  of this neutralized  $12.5 \text{ cm}^3$  of sodium hydroxide solution containing  $20 \text{ g/dm}^3$ . Determine the percentage purity of the calcium carbonate. [93.75%]
14. 0.15 gram of an ammonium salt was heated with excess sodium hydroxide. The ammonia given off was absorbed in  $25 \text{ cm}^3$  of  $0.16 \text{ M HCl}$ . The excess acid required  $12 \text{ cm}^3$  of  $0.2 \text{ M}$  sodium hydroxide for neutralization. Calculate the percentage of ammonia in the salt. [57%]
15.  $24 \text{ cm}^3$  of solution of tribasic acid containing 12 grams of the acid per  $\text{dm}^3$  of solution just neutralized  $48 \text{ cm}^3$  of  $0.09 \text{ M}$  sodium hydroxide. Calculate the molar mass of the acid. [200]
16. (a) (i) Explain the term primary standard solution (3marks)
- (ii) Why is sodium hydroxide NOT recommended as a primary standard for volumetric analysis? (2 marks)
- (b) (i) Why does the salt of sodium carbonate behaves as an alkali in acid-base titration?
- (ii)  $21.40 \text{ g}$  of hydrated sodium carbonate ( $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ ) was dissolved in water to make a litre of solution.  $25 \text{ cm}^3$  of this solution required  $18.70 \text{ cm}^3$  of  $0.2 \text{ M}$  standard hydrochloric acid solution for complete neutralization. Calculate the value of X. [10]
17. A hydrated compound Q contains the following ions, iron, sulphate and ammonium.
- (a) When  $9.64 \text{ g}$  of Q was heated,  $4.32 \text{ g}$  of water was evolved. Determine the number of moles of water of crystallization in Q. (1mark). [0.24moles]
- (b) When a solution containing  $9.64 \text{ g}$  of Q was heated strongly with excess sodium hydroxide solution,  $2.14 \text{ g}$  of hydroxide of iron was formed together with ammonia. The ammonia required  $100 \text{ cm}^3$  of  $0.2 \text{ M}$  hydrochloric acid for complete neutralization.
- Calculate

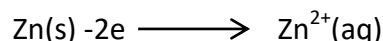
- (i) the number of moles of ammonium ions in Q (2marks) [0.02moles]
- (ii) the number of moles of irons in Q (2marks) [0.02moles]
- (c) A solution containing 9.64g of Q was treated with excess barium chloride solution gave 9.32g of barium sulphate. Determine the number of moles of sulphate ions in Q (2marks)[0.04moles]
- (d) Determine the molecular formula of Q. (2marks)  $[\text{Fe}(\text{NH}_4)(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}]$
18. Concentrated sulphuric acid is 18M. Calculate the volume, in  $\text{cm}^3$  of the acid that would be required to prepare  $200\text{cm}^3$  of 2.0M sulphuric acid (3 marks) [ $22.22\text{cm}^3$ ]
19. Concentrated nitric acid is 70% (w/w) and has a density of  $1.42\text{gcm}^{-3}$ .  $12.68\text{cm}^3$  of the acid in was dissolved in water and the solution made up to  $250\text{cm}^3$  with distilled water. Calculate the volume that would react completely with  $25.0\text{cm}^3$  of 0.2M sodium carbonate (4 marks) [ $12.5\text{cm}^3$ ]
20. 1.55g of an acid  $\text{C}_n\text{H}_{2n}(\text{COOH})_2$  was dissolved in water and the solution made up to  $250\text{cm}^3$ .  $25.0\text{cm}^3$  of this solution required  $23.5\text{cm}^3$  of 0.1M aqueous sodium hydroxide solution for complete neutralization. Calculate n (9 marks) [3]
21. 1.07g of a nitrogen containing compound was boiled with excess sodium hydroxide solution to produce ammonia. The ammonia produced neutralized  $200\text{cm}^3$  of a 0.1M monobasic acid. Calculate the percentage by mass of nitrogen in the compound [26.2%]

### Redox titration

Redox reaction is a reaction in which both reduction and oxidation take place simultaneously. A reduction reaction is where there is addition of electrons to an ion species;



whereas, oxidation is loss of electrons from a compound.



The species which accepts electrons such as  $\text{Cu}^{2+}$  is reduced but is referred to as an oxidizing agent. The species which donates electrons such as  $\text{Zn}^{2+}$  is oxidized but is referred to as a reducing agent

Oxidation Number

This is a number used to express the oxidation state of an element. Atoms of elements are

given oxidation number zero. Then when two elements are combined the atoms or ions of more electronegative element are regarded as being in negative oxidation state, and those of a more electropositive element in a positive state. If we arbitrary assume that the bonds between the atoms are all ionic, the oxidation number of an element is simply the number of electrons given up or received by an atom of an element. E.g. sodium chloride the oxidation number of sodium is 1, while that of chlorine is -1; hydrogen in HCl has oxidation number 1 but in sodium hydride, NaH is -1. Some elements have several oxidation states. Sulphur has oxidation number -2 in  $\text{H}_2\text{S}$ , 4 in  $\text{SO}_2$  and VI in  $\text{SO}_3$ . chlorine shows a wide variety of oxidation number, e.g. -1 in HCl, 0 in  $\text{Cl}_2$ , 1 in  $\text{ClO}^-$ , 3 in  $\text{ClF}_3$ , 4 in  $\text{ClO}_2$ , 5 in  $\text{ClO}_3^-$ , 6 in  $\text{ClO}_3$  and 7 in  $\text{ClO}_4^-$ .

An increase in oxidation number of an element during a reaction means that an element has been oxidized ; conversely, a decrease in

oxidation number means that an element has been reduced.

### Calculating oxidation number of ions of an element in complex ions.

Unless otherwise state or implied in these calculations, the oxidation number of a metal is equal to group number, e.g., Na is 1, Mg is 2 Al is 3. Oxidation number of O is -2 and that of any group 7 element is -1.

#### Example 34

Calculate the oxidation number of manganese ion in  $\text{KMnO}_4$ .

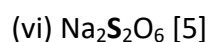
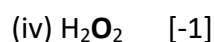
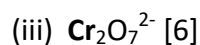
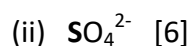
Here the oxidation number of K = 1 and that of oxygen is implied to be -2. if the oxidation number of Mn is X

Then,  $1 + X + -2 \times 4 = 0$  (overall charge ion the molecule)

$$X = 7$$

## Trial 7

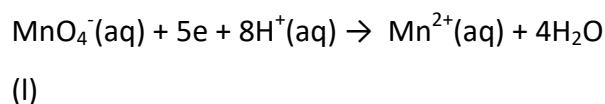
Calculate the oxidation numbers of the Bolden atoms in molecules or atoms below



### Half equations

These are balanced based on the number of electrons lost or accepted by a central atom.

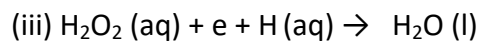
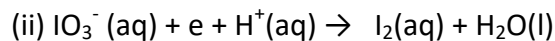
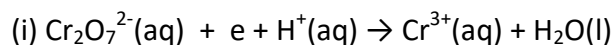
Example, equation for reduction of manganate (VII) in acidic medium is written as follows:



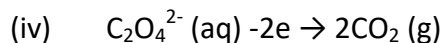
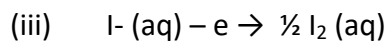
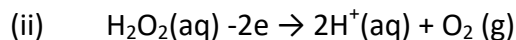
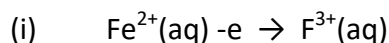
Five electrons are added because the oxidation number of Mn is reduced from 7 to 2.

### Exercise VI I

Balance the following reduction equations



The following are common oxidation half equation reactions:

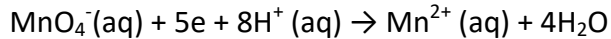


### Overall equation

These are constructed from half equation in such a way that the number electron received by an oxidizing agent is equal to the number of electron lost from a reducing agent.

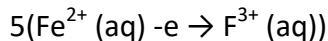
For example a reaction between  $\text{MnO}_4^-$  and  $\text{Fe}^{2+}$  in acidic medium, the half equation of  $\text{Fe}^{2+}$  is multiplied by 5 to provide the 5 electron necessary to reduce  $\text{MnO}_4^-$ .

That is,

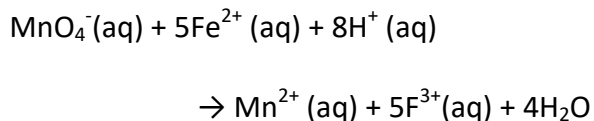


(l)

**Plus**



**Equals**



(l)

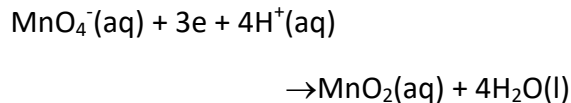
### Trial 8

- Write overall equations between the following pairs of ions or compounds, assume acidic medium where necessary
  - $\text{MnO}_4^-$  and  $\text{I}^-$
  - $\text{MnO}_4^-$  and  $\text{H}_2\text{O}_2$
  - $\text{KIO}_3$  and  $\text{KI}$
  - $\text{MnO}_4^-$  and  $\text{C}_2\text{O}_4^{2-}$
  - $\text{K}_2\text{Cr}_2\text{O}_7$  and  $\text{FeSO}_4$
- 0.9875g of an impure potassium manganate (VII) was dissolved in water to make  $250\text{cm}^3$  of solution. When  $20.0\text{cm}^3$  of this solution was acidified with dilute sulphuric acid, warmed and titrated against sodium ethanedioate (oxalate) solution, made by dissolving 1.675g of anhydrous sodium ethanedioate to make  $250\text{cm}^3$ ,

$24.40\text{cm}^3$  of sodium ethanedioate solution was used. ( $\text{Na}_2\text{C}_2\text{O}_4 = 134$  and  $\text{KMnO}_4 = 158$ ).

Calculate percentage purity of potassium permanganate (VII) [97.6%]

- $10\text{cm}^3$  of concentrated hydrogen peroxide was diluted to  $250\text{cm}^3$ .  $20\text{cm}^3$  of the dilute solution required  $28.30\text{cm}^3$  of  $0.02\text{M}$  potassium manganate (VII) solution for complete reaction. Find the volume strength of the concentrated hydrogen peroxide solution. [1.77M]
- Potassium permanganate is commonly used as a reagent in volumetric analysis. However it is not a primary standard.
  - State two disadvantages of using potassium manganate (VII) as a reagent in volumetric analysis.
  - Give two reasons why potassium manganate (VII) is not used as a primary standard
  - Write a half equation for the reduction of manganate (VII) ions in
    - acid solution
 
$$\text{MnO}_4^- (\text{aq}) + 5\text{e}^- + 8\text{H}^+ (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l})$$
    - alkaline solution



5. In volumetric estimation of reducing agent, potassium dichromate (VI) is preferred to potassium manganate (VII) as an oxidant

(a) Explain why potassium dichromate (VI) is preferred as an oxidant. (2 marks)

(b) 3.8g of solder containing tin was dissolved in excess hydrochloric acid. The solution was made up to 250cm<sup>3</sup>. 25.0cm<sup>3</sup> of this solution required 23.5cm<sup>3</sup> of a 0.01M potassium dichromate (VI) solution for complete reaction.

Calculate the percentage, by mass, of tin in the solder. [22%]

(c) Explain why chlorine is more soluble in dilute sodium hydroxide than in water (4 marks)

6. (a) Determine the oxidation number of

- i. Nitrogen in HNO<sub>3</sub> [5]
- ii. Sulphur in S<sub>2</sub>O<sub>8</sub><sup>2-</sup> (1marks) [8]

(b) Write half equation for the conversion of S<sub>2</sub>O<sub>8</sub><sup>2-</sup> to SO<sub>4</sub><sup>2-</sup> ions  

$$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$$

(c) Complete the following and balance the equations

- i.  $\text{S}_2\text{O}_8^{2-} + \text{I}^- \rightarrow$
- ii.  $\text{Sn}^{2+}(\text{aq}) + \text{S}_2\text{O}_8^{2-}(\text{aq}) \rightarrow$

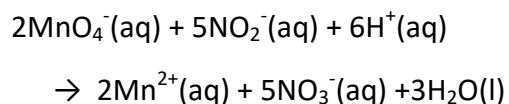
7. Iron (II) sulphate is normally used to standardize a solution of potassium

manganate (VII) acidified with sulphuric acid.

(i) write equation for the reaction between potassium manganate (VII) and iron (II) sulphate

(ii) State why hydrochloric acid is not used to acidify potassium manganate (VII) solution  
 Cl<sup>-</sup> are oxidized by MnO<sub>4</sub><sup>-</sup> as well

(b) 25 cm<sup>3</sup> of acidified solution of 0.02M potassium manganate (II) reacted exactly with 24.95 cm<sup>3</sup> of sodium nitrite. Potassium manganate (VII) reacts with sodium nitrite according to the following equation



Calculate the concentration of sodium nitrite in moles per litre. [0.05M]

## Iodometry

Here a substance is reacted with excess potassium iodide in redox equation to liberate iodine; the amount of iodine liberated is determined by titration with sodium thiosulphate solution to determine the concentration of the substance.

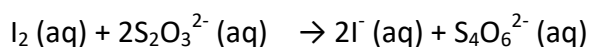
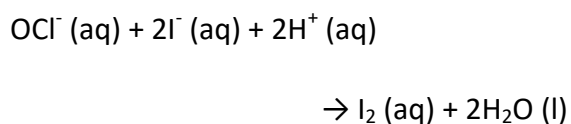
### Example

Find concentration of NaOCl in gram per litre given that 20cm<sup>3</sup> of NaOCl solution was reacted with excess potassium iodide, the liberated iodine required 25 cm<sup>3</sup> of 0.1M sodium thiosulphate solution.

### Solution

In acid medium, sodium chlorate reacts with iodide ions to liberate iodine, which oxidizes thiosulphate ions.

The reactions take place according to the following equations



Moles of sodium thiosulphate that reacted

100cm<sup>3</sup> contain 0.1moles

$$25\text{cm}^3 \text{ contain } \frac{0.1 \times 25}{100} = 0.0025\text{mols}$$

Moles of iodine produced from equation 2

$$= \frac{0.0025}{2} = 0.00125\text{moles}$$

### Moles NaOCl

From equation 1, moles of NaOCl is equal to the moles of iodine = 0.0015 moles

### Molarity of NaOCl

20cm<sup>3</sup> of NaOCl solution contain 0.00125 moles

$$1000\text{cm}^3 \text{ contain } \frac{0.00125 \times 1000}{20} = 0.0625\text{M}$$

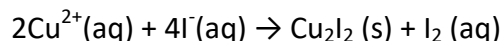
Concentration in grams per litre

$$\begin{aligned} 1 \text{ mole of NaOCl weighs} &= (40 + 16 + 35.5) \\ &= 91.5\text{g} \\ 0.0625 \text{ moles weigh} &= 91.5 \times 0.0625 \\ &= 5.7\text{g/l} \end{aligned}$$

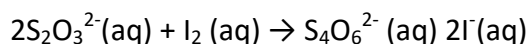
## Trial 9

### 1986/2/8e

- 6.53 g of impure copper was dissolved in excess nitric acid. The solution obtained mad up to 250cm<sup>3</sup> of solution with water. To 25cm<sup>3</sup> of this solution excess aqueous potassium iodide was added. The reaction that took place is given by the equation



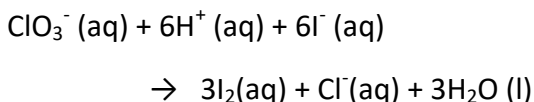
The iodine liberated reacted with 20cm<sup>3</sup> of 0.5M aqueous sodium thiosulphate according to the equation



Calculate the percentage of by mass the purity of copper in the sample (Cu =64) (6marks) [9.8%]

- 2.0g of a mixture of potassium chloride and potassium chlorate were dissolved in 250cm<sup>3</sup> of water. 10cm<sup>3</sup> of the solution was

mixed with excess potassium iodide. The iodine liberated required  $8.0\text{cm}^3$  of a  $0.2\text{M}$  sodium thiosulphate solution for complete reaction. Potassium chlorate and potassium chloride react according to the equation

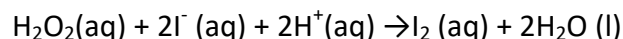


Calculate the percentage of potassium chlorate in the mixture (6marks) [1.6%]

3. During the determination of concentration of hydrogen peroxide in  $\text{g/dm}^3$  of the medical hydrogen peroxide,  $12\text{ cm}^3$  of medical hydrogen peroxide was diluted to  $250\text{cm}^3$  with distilled water.  $20\text{cm}^3$  of the

dilute solution was acidified required and excess KI added. The liberated iodine required  $22\text{cm}^3$  of sodium thiosulphate solution containing  $9.3\text{g Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$  in  $250\text{cm}^3$  of solution.

In acid medium, hydrogen peroxide reacts with iodide ions follows”



Calculate the volume strength of medical hydrogen peroxide. [20.625V]

### Double indicator titration.

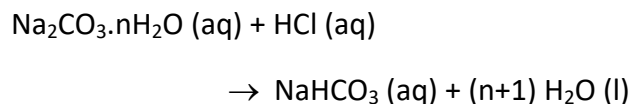
Titration requiring use of two indicators is employed to determine the concentrations and proportions of any two of the following base in a mixture, i.e.  $\text{NaOH}$ ,  $\text{Na}_2\text{CO}_3$  and  $\text{NaHCO}_3$ . Ideally, the two indicators referred to here are the phenolphthalein and methyl orange indicator. The basis of these titrations is that,  $\text{NaOH}$ , and  $\text{Na}_2\text{CO}_3$  are titrated to complete neutralization with methyl orange, whereas,  $\text{NaHCO}_3$  is acidic to phenolphthalein indicator and thus  $\text{Na}_2\text{CO}_3$  is only half neutralized when titrated with an acid in presence of phenolphthalein indicator.

#### Example 36

*7.15 g sodium carbonate  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$  was dissolved in water to make 250 ml of solution.  $25.0\text{ cm}^3$  of this solution required  $12.50\text{ cm}^3$  of  $0.2\text{ M HCl}$  using phenolphthalein indicator. Calculate the value if  $n$ .*

Equation

In presence of phenolphthalein indicator sodium carbonate is only half neutralized as per the equation



Moles of the acid that reacted:

$$\begin{aligned} 1000\text{ cm}^3 &\text{ contain } 0.2\text{ mole} \\ \therefore 12.5\text{ cm}^3 &\text{ contain } \frac{0.2 \times 12.5}{1000} \\ &= 0.0025\text{ mole} \end{aligned}$$

From the reaction equation, 1 mole of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$  reacts with 1 mole of  $\text{HCl}$ .

$\therefore$  0.0025 mole of the acid react with 0.0025 mole of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$ .

Molarity of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$

$\Rightarrow 25,0 \text{ cm}^3$  of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$  solution contain 0.0025 mole of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$ .

$$1000 \text{ cm}^3 \text{ contain } \frac{0.0025 \times 1000}{25} \\ = 0.1\text{M}$$

$\therefore$  Molarity of  $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$  solution is 0.1M

$$\text{RMM of } \text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O} = \frac{\text{Concentration g/l}}{\text{Molarity}} \\ = \frac{7.15 \times 4}{0.1} \\ = 286$$

$$\text{Thus, } (23 \times 2) + 12 + (16 \times 3) + n(2+16) = 286$$

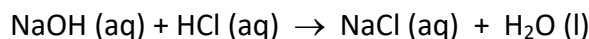
$$106 + 18n = 286$$

$$n = 10$$

### Example 35

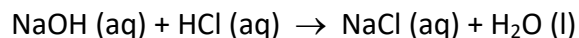
$25.0 \text{ cm}^3$  of a solution containing a mixture of sodium hydroxide and sodium hydrogen carbonate require  $15.0 \text{ cm}^3$  of 0.1M HCl using phenolphthalein indicator.  $25.0 \text{ cm}^3$  of the same mixture required  $28 \text{ cm}^3$  of the acid using methyl orange indicator. Calculate the proportion by mass of the two bases in the mixture.

Equation for the reaction that took place using phenolphthalein

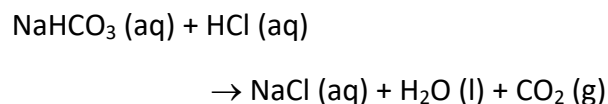


$\therefore 15 \text{ cm}^3$  of 0.1M HCl was required to neutralize NaOH is  $25 \text{ cm}^3$  of the mixture.

Equations using methyl orange indicator



and



$\therefore$  The difference between the volume ( $28 - 15 = 13 \text{ cm}^3$ ) of the acid was required to neutralize  $\text{NaHCO}_3$ .

No. of moles of the acid that reacted with NaOH

$$1000 \text{ cm}^3 \text{ contain } 0.1 \text{ mole}$$

$$\therefore 15 \text{ cm}^3 \text{ contain } \frac{0.1 \times 15}{1000} = 0.0015 \text{ mole}$$

From the reaction ratio of 1:1, 1 mole of HCl reacts with 1 mole of NaOH.

$$\Rightarrow \text{moles of NaOH in } 25 \text{ cm}^3 = 0.0015 \text{ mole}$$

$$1000 \text{ cm}^3 \text{ will contain } \frac{0.0015 \times 1000}{25} \\ = 0.06 \text{ M}$$

Concentration of NaOH in g/L

$$= \text{RMM} \times \text{Molarity} \\ = 40 \times 0.06 \\ = 2.4 \text{ g/dm}^3$$

No. of moles of the acid that reacted with  $\text{NaHCO}_3$ .

$$1000 \text{ cm}^3 \text{ contain } 0.1 \text{ mole}$$

$$\therefore 13 \text{ cm}^3 \text{ contain } \frac{0.1 \times 13}{1000} = 0.0013 \text{ mole}$$

From the reaction ratio of 1:1,

1mole of HCl reacts with 1 mole of the  $\text{NaHCO}_3$ .

⇒ moles of NaHCO<sub>3</sub> in 25 cm<sup>3</sup> = 0.0013 mole

$$1000 \text{ cm}^3 \text{ will contain } \frac{0.0013 \times 1000}{25} \\ = 0.052 \text{ M}$$

Concentration of NaHCO<sub>3</sub> in g/L

$$= \text{RMM} \times \text{Molarity} \\ = 84 \times 0.052 \\ = 4.368 \text{ g/dm}^3$$

The ratio of masses of NaOH : NaHCO<sub>3</sub> is

$$= 2.4 : 4.368 \\ = 1 : 1.82$$

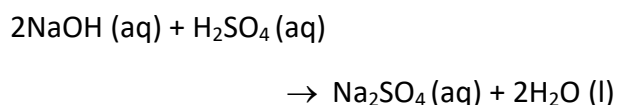
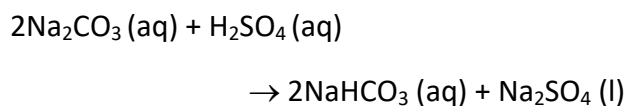
### Example 36

25.0 cm<sup>3</sup> of a solution containing a mixture of sodium carbonate and sodium hydroxide require 30.0 cm<sup>3</sup> of 0.1M H<sub>2</sub>SO<sub>4</sub> using phenolphthalein indicator. Another 25.0 cm<sup>3</sup> of the same mixture required 48.0 cm<sup>3</sup> of the acid using methyl orange indicator. Calculate the molar concentration of the mixture with respect to:

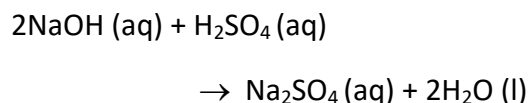
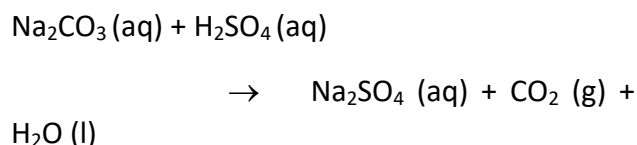
(i) Sodium carbonate (ii) sodium hydroxide. (9Marks)

Reaction Equations.

#### Using phenolphthalein indicator



#### Using methyl orange indicator



From the reactions above, the extra volume (48 - 30 = 18 cm<sup>3</sup>) of acid under methyl orange indicator is used to half neutralization of Na<sub>2</sub>CO<sub>3</sub>. This also is equal to the volume of the acid that converts Na<sub>2</sub>CO<sub>3</sub> to NaHCO<sub>3</sub>.

∴ the volume of the acid that completely neutralized Na<sub>2</sub>CO<sub>3</sub> in 25 cm<sup>3</sup> of solution = 18 x 2 = 36 cm<sup>3</sup>.

And volume of the acid that reacted

$$\text{with NaOH} \quad = 30 - 18 \\ = 12 \text{ cm}^3.$$

No. of moles of the acid that reacted with NaOH.

1000 cm<sup>3</sup> contain 0.1 mole

$$\therefore 12 \text{ cm}^3 \text{ contain } \frac{0.1 \times 12}{1000} = 0.0012 \text{ mole}$$

From the reaction ratio of 1:2, 1 mole of H<sub>2</sub>SO<sub>4</sub> reacts with 2 moles of NaOH.

$$\Rightarrow \text{moles of NaOH in } 25 \text{ cm}^3 \\ = 0.0012 \times 2 \text{ mole}$$

$$1000 \text{ cm}^3 \text{ will contain } \frac{0.0012 \times 2 \times 1000}{25}$$

$$= 0.048 \text{ M}$$

$$\therefore \text{Molarity of solution in respect to NaOH} \\ = 0.048 \text{ M.}$$

No. of moles of the acid that reacted with  $\text{Na}_2\text{CO}_3$ .

1000  $\text{cm}^3$  contain 0.1 mole

$$\therefore 36 \text{ cm}^3 \text{ contain } \frac{0.1 \times 36}{1000} = 0.0036 \text{ mole}$$

From the reaction ratio of 1:1, 1 mole of  $\text{H}_2\text{SO}_4$  reacts with 1 mole of the  $\text{Na}_2\text{CO}_3$ .

$$\Rightarrow \text{moles of } \text{Na}_2\text{CO}_3 \text{ in } 25 \text{ cm}^3 = 0.0036 \text{ mole}$$

$$\begin{aligned} 1000 \text{ cm}^3 \text{ will contain } & \frac{0.0036 \times 1000}{25} \\ & = 0.144 \text{ M} \end{aligned}$$

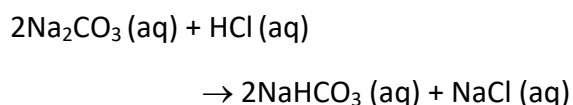
$\therefore$  the molarity of solution in respect to  $\text{Na}_2\text{CO}_3$  is **0.144M**

### Example 37

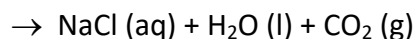
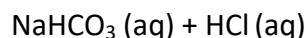
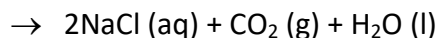
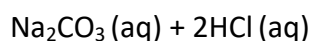
*25.0  $\text{cm}^3$  of a solution containing a mixture of sodium carbonate and sodium hydrogen carbonate require 15.0  $\text{cm}^3$  of 0.5M HCl using phenolphthalein indicator. Another 25.0  $\text{cm}^3$  of the same mixture required 34.5  $\text{cm}^3$  of the acid using methyl orange indicator. Calculate the concentration of the mixture with respect to: (i) sodium carbonate (ii) sodium hydrogen carbonate per  $\text{dm}^3$  of solution.*

Reaction Equations.

#### Using phenolphthalein indicator



#### Using methyl orange indicator



From the reactions above, 15.0  $\text{cm}^3$  of the acid was required for half neutralization of  $\text{Na}_2\text{CO}_3$ .

$\therefore$  the volume of the acid that completely neutralized  $\text{Na}_2\text{CO}_3$  in 25  $\text{cm}^3$  of solution  
 $= 15 \times 2 = 30 \text{ cm}^3$ .

and volume of the acid that reacted with  $\text{NaHCO}_3$  initially in 25.0  $\text{cm}^3$  of the mixture  
 $= 34.5 - 30$   
 $= 4.5 \text{ cm}^3$ .

No. of moles of the acid that reacted with  $\text{Na}_2\text{CO}_3$ .

1000  $\text{cm}^3$  contain 0.5 mole

$$\therefore 30 \text{ cm}^3 \text{ contain } \frac{0.5 \times 30}{1000} = 0.015 \text{ mole}$$

From the reaction ratio of 2:1, 2 mole of HCl react with 1 mole of the  $\text{Na}_2\text{CO}_3$

$$\begin{aligned} \Rightarrow \text{moles of NaOH in } 25 \text{ cm}^3 \\ = \frac{0.015}{2} = 0.0075 \text{ moles} \end{aligned}$$

$$\begin{aligned} 1000 \text{ cm}^3 \text{ will contain } & = \frac{0.0075 \times 1000}{25} \\ & = 0.3 \text{ M} \end{aligned}$$

Concentration of  $\text{Na}_2\text{CO}_3$  in g/L = RMM x Molarity

$$= 106 \times 0.3$$

$$= 31.8 \text{ g/dm}^3$$

No. of moles of the acid that reacted with  $\text{NaHCO}_3$

$1000 \text{ cm}^3$  contain 0.5 mole

$$\therefore 4.5 \text{ cm}^3 \text{ contain } \frac{0.5 \times 4.5}{1000} \times 0.5$$

$$= 0.00225 \text{ mole}$$

From the reaction ratio of 1:1, 1 mole of HCl reacts with 1 mole of  $\text{NaHCO}_3$ .

$$\Rightarrow \text{moles of } \text{NaHCO}_3 \text{ in } 25 \text{ cm}^3 \text{ _____}$$

$$= 0.00225 \text{ mole}$$

$$1000 \text{ cm}^3 \text{ will contain } \frac{0.00225 \times 1000}{25}$$

$$= 0.09 \text{ M}$$

Concentration of  $\text{NaHCO}_3$  in g/L

$$= \text{RMM} \times \text{Molarity}$$

$$= 84 \times 0.09$$

$$= \mathbf{7.56 \text{ g/dm}^3}$$

### Trial 10

1.  $25.0 \text{ cm}^3$  of a solution containing a mixture of sodium carbonate and sodium hydroxide require  $27.2 \text{ cm}^3$  of  $0.05\text{M}$   $\text{H}_2\text{SO}_4$  using phenolphthalein indicator. Another  $25.0 \text{ cm}^3$  of the same mixture required  $37.6 \text{ cm}^3$  of the acid using methyl orange indicator. Calculate the mass of each base in  $250 \text{ cm}^3$  of the mixture.

$$[\text{NaOH} = 0.672\text{g}, \text{Na}_2\text{CO}_3 = 1.1024\text{g}]$$

2.  $25.0 \text{ cm}^3$  of a solution containing a mixture of sodium carbonate and sodium hydrogen carbonate require  $11.2 \text{ cm}^3$  of  $0.1\text{M}$  HCl using phenolphthalein indicator and a further  $28.6 \text{ cm}^3$  of the acid using methyl orange

indicator. Calculate the percentage of sodium carbonate in the mixture. [69.51%]

3.  $25.0 \text{ cm}^3$  of a solution containing a mixture of sodium carbonate and sodium hydrogen carbonate required  $15.00 \text{ cm}^3$  of  $0.5\text{M}$  hydrochloric acid for complete reaction using phenolphthalein indicator.

$25.0 \text{ cm}^3$  of solution of the mixture required  $34.50 \text{ cm}^3$  of the acid using methyl orange indicator. Calculate the mass of sodium carbonate and sodium hydrogen carbonate in the solution in grams per litre. (6marks)

$$[\text{Na}_2\text{CO}_3 = 31.8\text{g}, \text{NaHCO}_3 = 7.56\text{g}]$$

**Thank you**  
**Dr. Bbosa Science**