

## Sub-topic 1.2: AMOUNT OF SUBSTANCE, MOLE CALCULATIONS.

### Mole

Define the term 'mole'

- A mole is the amount of a substance which contains  $6.02 \times 10^{23}$  particles.
- Or A mole is the amount of substance that contains the same number of particles as the number of particles in 12 grams of carbon-12.

The number  $6.02 \times 10^{23}$  is called Avogadro's number or constant and is denoted by letter  $L$

One mole of a substance has  $6.02 \times 10^{23}$  particles. The particles of substances grouped into moles can be molecules, atoms, ions, electrons, radicals, protons or any other specified particles. For example;

*1 mole of magnesium atoms contains  $6.02 \times 10^{23}$  atoms.*

*1 mole of magnesium ions contains  $6.02 \times 10^{23}$  ions.*

*1 mole of sulphuric acid molecules contains  $6.02 \times 10^{23}$  molecules of sulphuric acid*

### **MOLAR MASS OR FORMULA MASS OR MOLECULAR MASS**

This is the mass of one mole of a compound. It is obtained by adding the relative atomic masses of the atoms present in a compound. The formula mass of a compound is equal to the relative formula mass expressed in grams.

#### **Examples**

**Calculate the formula mass of the following compounds;**

a) Copper(II) sulphate crystals,  $CuSO_4 \cdot 5H_2O$

$$\begin{aligned} \text{Formula mass of } CuSO_4 \cdot 5H_2O &= 64 + 32 + (16 \times 4) + 5[(1 \times 2) + 16] \\ &= 250g \end{aligned}$$

b) Oxygen molecule,  $O_2$

$$\begin{aligned} \text{Formula mass of } O_2 &= 16 \times 2 \\ &= 32g \end{aligned}$$

c) Calcium dihydrogen phosphate,  $Ca(H_2PO_4)_2$

$$\begin{aligned} \text{Formula mass of } Ca(H_2PO_4)_2 &= 40 + (1 \times 2 \times 2) + (31 \times 2) + (16 \times 4 \times 2) \\ &= \end{aligned}$$

Calculate the relative formula masses of the following compounds;

- a)  $FeSO_4 \cdot 7H_2O$
- b)  $Al_2O_3$
- c)  $Ca(HCO_3)_2$
- d)  $KMnO_4$
- e)  $K_2Cr_2O_7$
- f)  $(NH_4)_2SO_4$

### **CALCULATING NUMBER OF PARTICLES**

This is based on the relationship that one mole of a substance contains  $6.02 \times 10^{23}$  particles.

- a) Calculate the number of particles in the following compounds.
  - i. Total number of ions in one mole of ammonium sulphate



*1 mole of ammonium sulphate contains 3 moles of ions.*

*1 mole of ion contains  $6.02 \times 10^{23}$  ions*

*3 moles of ions contain  $3 \times 6.02 \times 10^{23}$  ions*

$$= 18.06 \times 10^{23} \text{ ions}$$

- b) Calculate the number of atoms in the following;
  - i. 0.25 moles of calcium.
  - ii. 8 moles of sulphur.

### **EXERCISE**

1. Calculate the following;
  - a) Number of atoms in 2 moles of sodium.
2. Calculate the number of particles in the following;
  - a) 0.1 moles of sodium atoms.
  - b) 0.5 moles of chlorine atoms.
  - c) 0.3 moles of calcium atoms.
3. Use the value of  $6.02 \times 10^{23} \text{ mol}^{-1}$  for Avogadro's constant to find the number of atoms in.
  - a)  $2.0 \times 10^{-3} \text{ g}$  of calcium
  - b)  $5.0 \times 10^{-6} \text{ g}$  of argon

## CONVERTING NUMBER OF PARTICLES TO MASSES

### Example

1. Calculate the mass of sodium with  $1.5 \times 10^{22}$  sodium atoms.

$6.02 \times 10^{23}$  sodium atoms are contained in 1 mole of sodium.

$6.02 \times 10^{23}$  sodium atoms are contained in 23 g of sodium.

sodium atoms are contained in  $\left(\frac{23}{6.02 \times 10^{23}} \times 1.5 \times 10^{22}\right)$  g of sodium  
 $= 0.575$ g of sodium

2. How many grams of calcium contain.

a)  $6.02 \times 10^{23}$  atoms

$6.02 \times 10^{23}$  calcium atoms are contained in 1 mole of calcium.

$6.02 \times 10^{23}$  calcium atoms are contained in 40 g of calcium.

$6.02 \times 10^{23}$  calcium atoms are contained in  $\left(\frac{40}{6.02 \times 10^{23}} \times 6.02 \times 10^{23}\right)$  g of calcium  
 $= 40$  g of calcium

b)  $1.5 \times 10^{23}$  atoms

$6.02 \times 10^{23}$  atoms are contained in 1 mole of calcium.

$6.02 \times 10^{23}$  atoms are contained in 40 g of calcium.

$1.5 \times 10^{22}$  atoms are contained in  $\left(\frac{40}{6.02 \times 10^{23}} \times 1.5 \times 10^{22}\right)$  g of calcium  
 $= 10$  g of calcium.

c)  $3.0 \times 10^{23}$  atoms

.....

.....

## CONVERTING MASSES TO MOLES

### Relationships.

For atoms; 1 mole is equivalent to Relative atomic mass(RAM)in grams. Forexample;

1 mole of Cu weighs 64g; 1 mole of S weighs 32g;

For molecules; 1 mole is equivalent to Relative molecular mass(RMM) or Relative formula mass(RFM).Forexample; 1 mole of  $H_2O$  weighs 18g; 1 mole of  $CaCO_3$  weighs 100g;

1 mole of  $H_2SO_4$  weighs 98g

### Examples

1. Calculate the number of moles in;
  - a) 4g of oxygen molecule( $O = 16$ )  
 $RMM \text{ of oxygen, } O_2 = (2 \times 16)$   
 $= 32$   
 32 g contain 1 mole of oxygen.  
 4 g contain  $\left(\frac{1}{32} \times 4\right)$  moles of oxygen  
 $= 0.125$  moles of oxygen.
  - b) 160g of sulphur( $S = 32$ )  
 $RAM \text{ of sulphur} = 32$   
 32g is contained in 1 mole of sulphur.  
 160g is contained in  $\left(\frac{1}{32} \times 160\right)$  moles of sulphur.  
 $= 5.0$  moles of sulphur.
  
2. Calculate the number of moles in the following molecules.
  - a) 2g of calcium oxide.
  - b) 4g of sodium hydroxide
  - c) 22.2g of calcium chloride.( $Cl = 35.5$ )
  - d) 88g of carbon dioxide.( $C = 12, O = 16$ )
  - e) 9.8g of Sulphuric acid.
  - f) 45g of calcium dihydrogen phosphate.( $Ca = 40$ )
  - g) 197g of Disodium hydrogen phosphate( $Na = 23, H = 1, P = 31, S = 32$ )
  
3. Calculate the number of moles in;
  - a) 12g of magnesium( $Mg = 24$ )
  - b) 21.6g of silver( $Ag = 108$ )
  - c) 6g of ammonia ( $N = 14, H = 1$ )
  - d) 223g of carbon monoxide( $C = 12, O = 16$ )

#### **Summary**

$$\text{Number of moles} = \frac{\text{given mass}}{RAM}, \text{ for atoms.}$$

$$\text{Number of moles} = \frac{\text{given mass}}{RMM}, \text{ for Molecules/compounds.}$$

## CONVERTING MOLES TO MASSES

### Examples

Given; ( $O = 16, Ca = 40, C = 12, S = 32, H = 1, Cu = 64, Pb = 207, Cl = 35.5, N = 14$ )

1. Calculate the mass in;

a) 0.23 moles of sodium.

*1 mole of sodium weighs 23g*

$$\begin{aligned} 0.23 \text{ moles of sodium weigh } & \left( \frac{23}{1} \times 0.23 \right) g \\ & = 5.29g \text{ of sodium.} \end{aligned}$$

b) 7.1 moles of chlorine molecules.

*RMM of chlorine,  $Cl_2 = 35.5 \times 2 = 71$*

*1 mole of chlorine weighs 71g*

$$\begin{aligned} 7.1 \text{ moles of chlorine weigh } & \left( \frac{71}{1} \times 7.1 \right) g \\ & = 504.1g \text{ of chlorine.} \end{aligned}$$

c) 0.1 moles of potassium carbonate.

*RMM of potassium carbonate,  $K_2CO_3 = (39 \times 2) + 12 + (16 \times 3) = 138$*

*1 mole of potassium carbonate weighs 138g*

$$\begin{aligned} 0.1 \text{ mole of potassium carbonate weigh } & \left( \frac{138}{1} \times 0.1 \right) g \\ & = 13.8g \text{ of } K_2CO_3 \end{aligned}$$

d) 2.8 moles of nitrogen.

2. Calculate the mass of;

a) 0.1 mole of sodium atom.

b) 0.3 moles of chlorine molecules.

c) 0.05 moles of calcium atoms.

d) 0.05 moles of copper(II) carbonate.

e) 0.2 moles of lead(IV) oxide.

f) 0.02 moles of potassium manganate(VII)

## CALCULATIONS ON PERCENTAGE COMPOSITION BY MASS

From the formula of a compound, we can calculate the percentage by mass of each element in a compound as follows;

$$\text{Percentage composition of an element} = \left( \frac{\text{mass of an element}}{\text{Formula mass}} \times 100 \right) \%$$

### Examples

1. Calculate the percentage of nitrogen, oxygen and aluminium in aluminium nitrate.

(Al = 27; N = 14; O = 16)

$$\begin{aligned} \text{Molar mass of Al(NO}_3)_3 &= (27 \times 1) + (14 \times 3) + (16 \times 3 \times 3) \\ &= 27 + 42 + 144 \\ &= 213 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Mass of Aluminium} &= 27 \times 1 \\ &= 27 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Percentage of Aluminium} &= \left( \frac{27}{213} \times 100 \right) \\ &= 12.676056 \\ &\approx 12.7\% \end{aligned}$$

$$\begin{aligned} \text{Mass of Nitrogen} &= 14 \times 3 \\ &= 42 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Percentage of Nitrogen} &= \left( \frac{42}{213} \times 100 \right) \\ &= 19.718309 \\ &\approx 19.7\% \end{aligned}$$

$$\begin{aligned} \text{Mass of oxygen} &= 16 \times 3 \times 3 \\ &= 144 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Percentage of Oxygen} &= \left( \frac{144}{213} \times 100 \right) \\ &= 67.6056338 \\ &\approx 67.6\% \end{aligned}$$

2. Calculate the percentage of nitrogen in ammonium nitrate. (N = 14; H = 1; O = 16)
3. (a) Calcium dihydrogen phosphate is more used in agriculture as a source of phosphorus than calcium phosphate. (Ca = 40; H = 1; P = 31; O = 16)
- What general name can be given to the compounds in (a) that supplies plant nutrients to the soil.
  - Write down the chemical formulae of
    - Calcium dihydrogen phosphate.
    - Calcium phosphate.

- (b) Calculate the percentage of phosphorus in;
- Calcium dihydrogen phosphate.
  - Calcium phosphate.
- (c) Suggest a reason why calcium dihydrogen phosphate is more suitable in Agriculture than calcium phosphate.
- (d) Write an ionic equation to show how calcium phosphate can be prepared in the laboratory
4. Ammonium nitrate and ammonium sulphate are both used in Agriculture as source of nitrogen to the plants in the soil. ( $N = 14$ ;  $H = 1$ ;  $O = 16$ ;  $S = 32$ )
- Write down the chemical formulae of;
    - Ammonium nitrate.
    - Ammonium sulphate.
  - Calculate the percentage of nitrogen in;
    - Ammonium nitrate.
    - Ammonium sulphate.
  - Which of the two nitrogen supplying fertilizers is the best plant nutrient to the soil? Give a reason for your answer.
  - Write equation for the reaction to show how ammonium sulphate can be prepared in the laboratory.
 
$$\checkmark \text{NH}_3(g) + \text{H}_2\text{SO}_4(l) \rightarrow (\text{NH}_4)_2\text{SO}_4(s)$$
5. Calculate the percentage of;
- Calcium in calcium carbonate. ( $Ca = 40$ ;  $C = 12$ ;  $O = 16$ )
  - Oxygen in Triiron tetra oxide. ( $Fe = 56$ ;  $O = 16$ )
  - Oxygen in baking powder,  $\text{NaHCO}_3$  ( $Na = 23$ ;  $H = 1$ ;  $C = 12$ ;  $O = 16$ )
  - Nitrogen in sodium nitrate. ( $Na = 23$ ;  $N = 14$ ;  $O = 16$ )

6.

***CALCULATING PERCENTAGE OF WATER OF CRYSTALLISATION***

- a) Calculate the percentage of water of crystallization in the following compounds;

- $$\text{CuSO}_4 \cdot 5\text{H}_2\text{O} \quad (\text{Cu} = 64; \text{S} = 32; \text{O} = 16)$$

$$\text{Molar mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 64 + 32 + 64 + 90$$

$$= 250 \text{ g}$$

$$\text{Percentage of water of crystallisation} = \left( \frac{90}{250} \times 100 \right)$$

$$= 36\%$$

- $$\text{MgSO}_4 \cdot 7\text{H}_2\text{O} \quad (\text{Mg} = 24)$$

iii.  $Na_2CO_3 \cdot 10H_2O$

- b) Calculate the value of  $x$  in hydrated aluminium sulphate,  $Al_2(SO_4)_3 \cdot xH_2O$  if the percentage of aluminium is 14.3%. ( $Al = 27; S = 32; O = 16; H = 1$ )

Let the molar mass of  $Al_2(SO_4)_3 \cdot xH_2O$  be  $y$

Percentage of aluminium in  $Al_2(SO_4)_3 \cdot xH_2O$

$$= \left( \frac{\text{mass of aluminium}}{\text{molar mass of } Al_2(SO_4)_3 \cdot xH_2O} \times 100 \right)$$
$$14.3 = \frac{(27 \times 2)}{y} \times 100$$

$$14.3 = \frac{54}{y} \times 100$$

$$14.3y = 5400$$

$$y = 377.6 \text{ g}$$

$\therefore$  The Molar mass of  $Al_2(SO_4)_3 \cdot xH_2O = 377.6 \text{ g}$

$$Al_2(SO_4)_3 \cdot xH_2O = 377.6$$

$$(27 \times 2) + (32 \times 3) + (16 \times 4 \times 3) + x[(1 \times 2) + (16 \times 1)] = 377.6$$

$$54 + 96 + 193 + 2x + 16x = 377.6$$

$$342 + 18x = 377.6$$

$$18x = 377.6 - 342$$

$$18x = 35.6$$

$$x = 1.9778$$

$$x \approx 2$$

**Therefore, the value of  $x$  is 2**

- c) Hydrated sodium carbonate,  $Na_2CO_3 \cdot nH_2O$  contains 16.1% sodium. Calculate the value of  $n$  hence write the formula of the compound.

$$(Na = 23; C = 12; O = 16; H = 1)$$

- d) Hydrated iron(II) sulphate,  $FeSO_4 \cdot yH_2O$  contains 20.1% of iron. Determine the value of  $y$  ( $Fe = 56; S = 32; O = 16; H = 1$ )

- e) A metal sulphate,  $X_2(SO_4)_3$ , contains 28% by mass of metal X. determine the relative atomic mass of X and the relative molecular mass of  $X_2(SO_4)_3$

$$(S = 32; O = 16)$$

$$\text{(Answer; RAM of X = 56; RMM of } X_2(SO_4)_3 = 400)$$

## **EMPIRICAL FORMULA AND MOLECULAR FORMULA**

### **EMPIRICAL FORMULA**

This is the simplest formula of a compound which expresses the ratio in which different atoms present in one molecule exists.

Or

This is the formula that shows the simplest ratio of the different atoms in a molecule (or compound).

Empirical formula is sometimes referred to as simplest formula.

The empirical formula of a compound can be determined if the composition by mass of each element present in the compound is known. The mass composition may be given in grams or percentages.

### **Steps involved in calculating Empirical formulae**

- Write down the symbols of the elements present.
- Write down the percentage composition or composition by mass below the symbols.
- Find the number of moles of each element by dividing the percentage composition or mass by Relative Atomic mass of each element.
- Find the mole ratio of the elements by dividing the moles with the smallest number of moles.
- Then write down the empirical number
- If the mole ratio is in fractions; then the simplest ratio is determined as follows;
  - ✓ Round off to the nearest whole number if it is very close to the whole number.
  - ✓ Multiply by a small number that converts the fraction to a whole number if the fraction is not close to a whole number. E.g. if the mole ratio is (0.5, 1.5, 2.5, 3.5.....etc) multiply by 2;  
(0.33333..., 0.6666....., 1.3333....., 1.6666....., .....etc) multiply 3;  
(0.25,.....etc) multiply by 4.

### **MOLECULAR FORMULA**

This is the formula that shows the actual number of each atoms present in one molecule of a compound.

The molecular formula of a compound is a multiple of the empirical formula, so, from the empirical formula, molecular formula can be determined;

*Molecular formula = (Empirical formula)<sub>n</sub> = Molecular mass.*

where  $n$  is the number to be determined.

### **CALCULATIONS OF EMPIRICAL AND MOLECULAR FORMULAE**

#### **Examples;**

1. An organic compound contains 75% by mass of carbon and the rest is hydrogen.

Given that the molecular mass of the compound is 16. Calculate the;

- i. Empirical formula of the compound. ( $C = 12; H = 1$ )

$$\begin{aligned}\text{Percentage of Hydrogen} &= 100 - 75 \\ &= 25\%\end{aligned}$$

<i>Elements present</i>	<i>C</i>	<i>H</i>
<i>Percentage composition</i>	75	25
<i>Moles of atoms</i>	$\frac{75}{12}$ $= 6.25$	$\frac{25}{1}$ $= 25$
<i>Mole ratio</i>	$\frac{6.25}{6.25}$ $= 1$	$\frac{25}{6.25}$ $= 4$
<i>Simplest ratio</i>	1	: 4

**$\therefore$  The empirical formula of the compound is  $CH_4$**

- ii. The molecular formula of the compound.

$$\text{Molecular formula} = (\text{Empirical formula})_n = \text{Molecular mass.}$$

$$(\text{Empirical formula})_n = \text{Molecular mass.}$$

$$(CH_4)_n = 16.$$

$$(12 \times 1 + 1 \times 4)_n = 16$$

$$16n = 16$$

$$n = 1$$

**$\therefore$  The molecular formula of the compound is  $CH_4$**

*N.B. Sometimes, molecular mass is given in form of vapour density.*

*Therefore, it is calculated as follows;*

$$\text{Molecular mass} = 2 \times \text{vapour density.}$$

*Hence;*

$$(\text{Empirical formula})_n = 2 \times \text{vapour density}$$

2. A hydrocarbon Q contains 82.8% carbon. Calculate;
- The empirical formula of Q. ( $C = 12; H = 1$ )  

$$\text{Percentage of Hydrogen} = 100 - 82.8$$

$$= 17.2\%$$

<i>Elements present</i>	<i>C</i>	<i>H</i>
<i>Percentage composition</i>	82.8	17.2
<i>Moles of atoms</i>	$\frac{82.8}{12}$ = 6.9	$\frac{17.2}{1}$ = 17.2
<i>Mole ratio</i>	$\frac{6.9}{6.9}$ = 1	$\frac{17.2}{6.9}$ = 2.5
<i>Simplest ratio</i>	(1 × 2) 2	: : (2.5 × 2) 5

**∴ The empirical formula of Q is  $C_2H_5$**

- The molecular formula of Q. (Vapour density of Q = 29)  

$$(\text{Empirical formula})_n = 2 \times \text{vapour density}$$

$$(C_2H_5)_n = 2 \times 29$$

$$[(12 \times 2) + (1 \times 5)]n = 58$$

$$24n + 5n = 58$$

$$29n = 58$$

$$n = 2$$

**Therefore, molecular formula of Q is  $C_4H_{10}$**

- 6.71g of an oxide of tin contains 1.42g of oxygen. Calculate the empirical formula of the tin oxide. ( $Sn = 119; O = 16$ )
- A white powder is made up of 24% carbon and the 76% fluorine.  
 $(C = 12; F = 19)$

Calculate the molecular formula of the compound if the vapour density is 25.

- A compound R with molecular mass 60 consists of carbon 40%, hydrogen 6.7% and the rest being oxygen. Calculate the molecular formula of R.
- 1.5g of hydrocarbon M consists of 1.2g of carbon.
  - Calculate the empirical formula of M. ( $C = 12; H = 1$ )
  - Determine the molecular formula of M if its vapour density is 15.
- 1.66g of an oxide of metal X gave 1.18g of X on reduction. Determine the simplest formula of the oxide. ( $X = 59; O = 16$ )

8. When 6.4g of an oxide of element Y was heated and hydrogen passed over it. 3.2g of the element was produced. Determine the empirical formula of the oxide.  
(Y = 32; O = 16)
9. When 9.8g of an organic compound M containing carbon and hydrogen was burnt in excess oxygen, 31.55g of carbon dioxide and 10.76g of water was formed. Calculate the;

- i. Empirical formula of M. (C = 12; H = 1)
- Molar mass of  $CO_2 = (12 \times 1) + (16 \times 2) = 44 \text{ g}$   
 44g of carbon dioxide contain 12 g of carbon.  
 31.55g of carbon dioxide contain  $\left(\frac{12}{44} \times 31.55\right) \text{ g of carbon}$   
 $= 8.6045 \text{ g of carbon.}$   
 $\approx 8.60 \text{ g of carbon.}$
- Molar mass of  $H_2O = (1 \times 2) + (16 \times 1)$   
 $= 18 \text{ g}$   
 18 g of water contain 2g of hydrogen.  
 10.76 g of water contain  $\left(\frac{2}{18} \times 10.76\right) \text{ g of hydrogen.}$   
 $= 1.1955 \text{ g of hydrogen.}$   
 $\approx 1.20 \text{ g of hydrogen.}$

<i>Elements present</i>	<i>C</i>	<i>H</i>
<i>Composition by mass</i>	8.60	1.20
<i>Moles of atoms</i>	$\frac{8.60}{12}$ $= 0.7167$	$\frac{1.20}{1}$ $= 1.20$
<i>Mole ratio</i>	$\frac{0.7167}{0.7167}$ $= 1$	$\frac{1.20}{0.7167}$ $= 1.6743$
<i>Simplest ratio</i>	1	: 2

**$\therefore$  The empirical formula of M is  $CH_2$**

- ii. Molecular formula of M. (Vapour density of M = 12)
- $(\text{Empirical formula})_n = 2 \times \text{vapour density}$   
 $(CH_2)_n = 2 \times 12$   
 $[(12 \times 1) + (1 \times 2)]n = 24$   
 $14n = 24$   
 $n = 1.7$   
 $n \approx 2$

**$\therefore$  The molecular formula of M is  $C_2H_4$**

10. 3.10g of an organic compound that contains carbon, hydrogen and oxygen only produced 4.40g of carbon dioxide and 2.70g of water on complete combustion.

Calculate the;

- i. Empirical formula of the compound. ( $C = 12; H = 1; O = 16$ )
- ii. Molecular formula of the compound if its molecular mass=62g.

11. A gaseous hydrocarbon P contains 92.3% by mass of carbon. ( $C = 12; H = 1$ )

a) Determine the empirical formula of P.

$$\begin{aligned} \text{Percentage of hydrogen} &= 100 - 92.3 \\ &= 7.7\% \end{aligned}$$

<i>Elements present</i>	<i>C</i>	<i>H</i>
<i>Composition by mass</i>	92.3	7.7
<i>Moles of atoms</i>	$\frac{92.3}{12}$ = 7.7	$\frac{7.7}{1}$ = 7.7
<i>Mole ratio</i>	$\frac{7.7}{7.7}$ = 1	$\frac{7.7}{7.7}$ = 1
<i>Simplest ratio</i>	1	: 1

**$\therefore$  The empirical formula of P is CH**

b) If 0.13g of P occupies 112cm<sup>3</sup> at s.t.p. calculate the molecular formula of P  
(1 mole of a gas occupies 22400cm<sup>3</sup> at s.t.p)

112cm<sup>3</sup> of P has a mass of 0.13g

$$\begin{aligned} 22400\text{cm}^3 \text{ of P has a mass of } &\left(\frac{0.13}{112} \times 22400\right) \text{ g} \\ &= 26 \text{ g} \end{aligned}$$

Molecular mass of P = 26 g

$$(\text{Empirical formula})_n = \text{molecular mass}$$

$$(\text{CH}_2)_n = 26$$

$$[(12 \times 1) + (1 \times 1)]n = 26$$

$$13n = 26$$

$$n = 2$$

**$\therefore$  The molecular formula of P is C<sub>2</sub>H<sub>2</sub>**

12. 1.00g of Z contains 0.86g of carbon and the rest is hydrogen.

(1 mole of Z weighs 28g) ( $C = 12; H = 1$ )

Determine the molecular formula of Z.

13. A gaseous hydrocarbon X contains 14.3% hydrogen by mass. 6.0g of X occupy 4.8dm<sup>3</sup> at s.t.p. calculate the;
- Empirical formula of X. ( $C = 12; H = 1$ )
  - The molar mass of X. (*1 mole of a gas occupies 22400cm<sup>3</sup> at s. t. p*)
  - The molecular formula of X
14. 1.5g of a hydrocarbon Y contains 0.3g of hydrogen.
- Calculate the empirical formula of Y. ( $C = 12; H = 1$ )
  - 0.125g of hydrocarbon Y occupies a volume of 100cm<sup>3</sup> at room temperature and pressure. (*1 mole of gas occupies 24000cm<sup>3</sup> at r. t. p*)
    - Calculate the relative formula mass of Y
    - Determine the molecular formula of Y hence draw the structure of Y
15. When excess carbon monoxide was passed over 4.0g of a heated oxide of iron, Y, 2.8g of iron was formed. Determine the molecular formula of Y.  
( $O = 16; Fe = 56; Y = 160$ )

### **WATER OF CRYSTALLISATION.**

Define the term 'crystal'

- A crystal is a solid that consists of particles (atoms, molecules or ions) arranged in an orderly and repetitive manner resulting into a definite shape.

Or

- A crystal is a solid or substance that has solidified in a definite regular shape.

Define the term 'water of crystallization'

- This is a definite amount of water with which some substances chemically combine when they form crystals from their aqueous solutions.

Substances that contain water of crystallization are called 'crystals' or 'hydrated substances'

#### **Examples of hydrated salts/substances**

<b>Chemical name</b>	<b>Chemical formula</b>
<i>Copper(II)sulphate – 5 – water</i>	$CuSO_4 \cdot 5H_2O$
<i>Sodium carbonate – 10 – water</i>	$Na_2CO_3 \cdot 10H_2O$

Sodium sulphate – 10 – water	$Na_2SO_4 \cdot 10H_2O$
Magnesium sulphate – 7 – water	$MgSO_4 \cdot 7H_2O$
Iron(II)sulphate – 7 – water	$FeSO_4 \cdot 7H_2O$
Oxalic acid – 2 – water	$H_2C_2O_4 \cdot 2H_2O$

**Determination of water of crystallisation in a substance.**

- 4.10g of magnesium sulphate,  $MgSO_4 \cdot XH_2O$ , gave 2.0g of the anhydrous salt on heating. Calculate the value of X and hence deduce the molecular formula of the hydrated salt.

$$\text{Mass of water of crystallisation} = 4.1 - 2.0 = 2.1g$$

$$\begin{aligned} \text{Molar mass of water, } H_2O &= (1 \times 2) + (16 \times 1) \\ &= 18g \end{aligned}$$

$$\text{Mass of magnesium sulphate (anhydrous)} = 2.0g$$

$$\text{Molar mass of } MgSO_4 = (24 \times 1) + (32 \times 1) + (16 \times 4) = 120g$$

Compounds present	$MgSO_4$	$H_2O$
Composition by mass	2.0	2.1
Number of moles	$\frac{2.0}{120}$ $= 0.0167$	$\frac{2.1}{18}$ $= 0.1167$
Mole ratio	$\frac{0.0167}{0.0167}$ $= 1$	$\frac{0.1167}{0.0167}$ $= 6.988$
Simplest ratio	1	7

**$\therefore X = 7$  and the formula of the salt is  $MgSO_4 \cdot 7H_2O$**

**ALTERNATIVELY;**

$$\text{Mass of water of crystallisation} = 4.1 - 2.0 = 2.1g$$

$$\begin{aligned} \text{Molar mass of water, } H_2O &= (1 \times 2) + (16 \times 1) \\ &= 18g \end{aligned}$$

$$\begin{aligned} \text{Number of moles of water} &= \frac{2.1}{18} \\ &= 0.1167 \end{aligned}$$

$$\text{Mass of magnesium sulphate (anhydrous)} = 2.0g$$

Molar mass of  $MgSO_4 = (24 \times 1) + (32 \times 1) + (16 \times 4) = 120g$

$$\begin{aligned} \text{Number of moles of magnesium sulphate} &= \frac{2.0}{120} \\ &= 0.0167 \end{aligned}$$

0.0167 moles of magnesium sulphate combined with 0.1167 moles of water.

1 mole of magnesium sulphate will combine with  $\left(\frac{0.1167}{0.0167} \times 1\right)$  moles of water.  
 $= 6.988$  moles of water.  
 $\approx 7$  moles of water.

**$\therefore$  The value of  $X = 7$  and the formula of the salt is  $MgSO_4 \cdot 7H_2O$**

- When 5.0g of hydrated copper(II) sulphate were heated to a constant mass of 3.2g of anhydrous salt remained. Determine the formula of the hydrated salt.  
( $Cu = 64; S = 32; O = 16; H = 1$ )
- When 5.72g of a hydrated salt of sodium carbonate were heated to a constant mass 2.12g of the anhydrous salt remained. If the formula of the compound is  $Na_2CO_3 \cdot XH_2O$ . calculate the value of X. ( $Na = 23; C = 12; O = 16; H = 1$ )
- When 14.4g of hydrated sodium carbonate,  $Na_2CO_3 \cdot nH_2O$  was strongly heated, the mass of the residue was 5.4g.
  - Write the equation for the reaction that took place.
  - Calculate the;
    - Number of moles of water of crystallization. ( $H = 1; O = 16$ )
    - The percentage of water of crystallization.
- A hydrated salt T, consists of 20.2% iron, 11.5% Sulphur, 23% oxygen and 4.3% water of crystallization.
  - Calculate the empirical formula of T.  
( $Fe = 56; S = 32; O = 16; H = 1$ )
  - Deduce the molecular formula of T. (1 mole of T weighs 278g)
  - Write a molecular and ionic equation for the reaction between a solution of T and chlorine.
    - $6FeSO_4(aq) + 3Cl_2(g) \rightarrow 2FeCl_3(aq) + 2Fe_2(SO_4)_3(aq)$
    - $2Fe^{2+}(aq) + Cl_2(g) \rightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$
- When sodium carbonate crystals  $Na_2CO_3 \cdot XH_2O$  were heated strongly and there was a loss in mass of 62.9%. What is the formula of the crystal?  
( $Na = 23; C = 12; O = 16; H = 1$ )

$$\text{Percentage of Na}_2\text{CO}_3 = 100 - 62.9 = 37.1\%$$

$$\text{Molar mass of Na}_2\text{CO}_3 = (23 \times 2) + (12 \times 1) + (16 \times 3) = 106 \text{ g}$$

$$\text{Molar mass of H}_2\text{O} = (1 \times 2) + (16 \times 1) = 18 \text{ g}$$

<i>Compounds present</i>	$\text{Na}_2\text{CO}_3$	$\text{H}_2\text{O}$
<i>Percentage composition</i>	37.1	62.9
<i>Number of moles</i>	$\frac{37.1}{106}$ = 0.35	$\frac{62.9}{18}$ = 3.49
<i>Mole ratio</i>	$\frac{0.35}{0.35}$ = 1	$\frac{3.49}{0.35}$ = 9.98
<i>Simplest ratio</i>	1 : 10	

$\therefore X = 10$  and the formula of the crystal is  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

- Hydrated iron(II) sulphate,  $\text{FeSO}_4 \cdot n\text{H}_2\text{O}$  contains 20.1% of iron. Calculate the value of n. ( $\text{Fe} = 56; \text{S} = 32; \text{O} = 16; \text{H} = 1$ )
- A compound Z of molecular formula  $\text{A}_x\text{B}_y \cdot n\text{H}_2\text{O}$  consists of 8.57% A, 45.71% B and 45.72% of water.
  - Determine the values of x, y and n. ( $\text{A} = 27; \text{B} = 96; \text{H} = 1; \text{O} = 16$ )
  - Deduce the molecular formula of Z.
- When 7.2g of hydrated sodium carbonate ( $\text{Na}_2\text{CO}_3 \cdot n\text{H}_2\text{O}$ ) was strongly heated, the mass of the residue was 2.7g.
  - Calculate the number of moles of water of crystallization (value of n) ( $\text{Na} = 23; \text{C} = 12; \text{O} = 16; \text{H} = 1$ )
- In an experiment with hydrated magnesium sulphate,  $\text{MgSO}_4 \cdot X\text{H}_2\text{O}$ . The following results were obtained.
 

Mass of crucible + lid = 14.636g  
 Mass of crucible + lid + hydrated magnesium sulphate = 15.374g  
 Mass of crucible + lid + crystals after heating = 14.996g  
 Mass of crucible + lid + crystals after further heating = 14.996g

  - Determine the mass of water of crystallization in the hydrated magnesium sulphate.
  - What was the mass of anhydrous magnesium sulphate formed?

c) Determine the value of X. ( $Mg = 24; S = 32; O = 16; H = 1$ )

(Answer;  $X = 8$ )

d) Suggest why the last step was necessary

### **CALCULATING FORMULA MASS OR MOLECULAR MASS OR MOLAR MASS USING MOLAR GAS VOLUMES**

Define the term 'molar gas volume'

This refers to the volume occupied by 1 mole of a gas at standard conditions of temperature and pressure.

i.e.

1 mole of a gas occupies 22.4 l at s. t. p (standard temperature and pressure)

1 mole of a gas occupies 22.4dm<sup>3</sup> at s. t. p

1 mole of a gas occupies 22400cm<sup>3</sup> at s. t. p

1 mole of a gas occupies 24 l at r. t. p (room temperature and pressure)

1 mole of a gas occupies 24dm<sup>3</sup> at r. t. p

1 mole of a gas occupies 24000cm<sup>3</sup> at r. t. p

N.B. All the above are molar gas volumes at both s.t.p and r.t.p.

1. 350cm<sup>3</sup> of a certain gas weighed 1.5g of at s.t.p. calculate the molar mass of the gas.

(1 mole of a gas occupies 22.4 l at s. t. p)

350cm<sup>3</sup> of the gas weigh 1.5 g

$$22400\text{cm}^3 \text{ of the gas weigh } \left( \frac{1.5}{350} \times 22400 \right) \text{ g} \\ = 96 \text{ g}$$

**$\therefore$  The molar mass of the gas = 96 g**

2. Calculate the molar mass of gas Z if 15dm<sup>3</sup> of the gas weigh 7g at room temperature and pressure. (molar gas volume at r. t. p = 24dm<sup>3</sup>)

15dm<sup>3</sup> of the gas weigh 7 g

$$24\text{dm}^3 \text{ of the gas weigh } \left( \frac{7}{15} \times 24 \right) \text{ g} \\ = 11.2 \text{ g}$$

**$\therefore$  The molar mass of gas Z = 11.2 g**

3. 7.5g of a compound X occupy 5.6dm<sup>3</sup> at s.t.p. determine the molecular mass of X.

(molar gas volume at s. t. p = 22.4dm<sup>3</sup>)

### **CALCULATING VOLUMES OF A GAS**

1. The molar mass of gas M is 44g. calculate the volume occupied by 23g of the gas at

s.t.p. (1 mole of M occupies 22.4dm<sup>3</sup>)

1 mole of M occupies 22.4dm<sup>3</sup> at s. t. p

44g of M contain 22.4dm<sup>3</sup>

$$\begin{aligned}
 23\text{g of } M \text{ contain } & \left(\frac{22.4}{44} \times 23\right) \text{ dm}^3 \\
 & = 11.709 \text{ dm}^3 \text{ of } M \\
 & \approx \mathbf{11.7 \text{ dm}^3 \text{ of } M}
 \end{aligned}$$

**ALTERNATIVELY;**

**Moles of M**

44g of M contain 1 mole

$$\begin{aligned}
 23\text{g of } M \text{ contain } & \left(\frac{1}{44} \times 23\right) \text{ mole} \\
 & = \frac{23}{44} \text{ moles of } M
 \end{aligned}$$

But 1 mole of M occupies 22.4 dm<sup>3</sup> at s. t. p

$$\begin{aligned}
 \frac{23}{44} \text{ moles of } M \text{ occupy } & \left(\frac{22.4}{1} \times \frac{23}{44}\right) \text{ dm}^3 \\
 & = 11.709 \text{ dm}^3 \text{ of } M \\
 & \approx \mathbf{11.7 \text{ dm}^3 \text{ of } M}
 \end{aligned}$$

**$\therefore$  The volume of M = 11.7 dm<sup>3</sup>**

- Calculate the volume occupied by 50g of a gas whose formula mass is 32g at room temperature. (molar gas volume = 24 dm<sup>3</sup> at r. t)

### ***CALCULATING MASSES AND VOLUMES FROM STOICHIOMETRIC EQUATIONS.***

A stoichiometric equation is one in which the reactants and products are correctly balanced. Such an equation gives correct mole ratios of products and reactants in a given reaction. Since equations are quantitative statements, it is possible to calculate the exact quantities of products and reactants involved.

#### ***Calculating masses from stoichiometric equations***

- Calculate the loss in mass when 10g of calcium carbonate are heated to a constant mass. (Ca = 40; C = 12; O = 16)
- Copper(II) oxide is reduced by hydrogen according to the following equation.



Calculate the mass of copper formed when 4g of the oxide is reduced.

(Cu = 64; O = 16; H = 1)

Molar mass of CuO = (64 × 1) + (16 × 1) = 80g

From the equation;

1 mole of CuO produced 1 mole of copper

80g of copper(II) oxide produced 64g of copper.

4g of copper(II) oxide produce  $\left(\frac{64}{80} \times 4\right)$  g of copper.

$$= 3.2 \text{ g of copper.}$$

**∴ The mass of copper formed is 3.2 g**

**ALTERNATIVELY;**

**Moles of copper(II)oxide that reacted**

$$\text{Molar mass of CuO} = (64 \times 1) + (16 \times 1) = 80\text{g}$$

80 g of copper(II)oxide contain 1 mole.

$$4 \text{ g of copper(II)oxide contain } \left(\frac{1}{80} \times 4\right) \text{ moles}$$

$$= 0.05 \text{ moles of copper(II)oxide.}$$

**Moles of copper formed.**

From the equation;

1 mole of copper(II)oxide produced 1 mole of copper

$$0.05 \text{ moles of copper(II)oxide produced } \left(\frac{1}{1} \times 0.05\right) \text{ moles of copper}$$

$$= 0.05 \text{ moles of copper.}$$

**Mass of copper formed**

1 mole of copper weighs 64 g

$$0.05 \text{ moles of copper will weigh } \left(\frac{64}{1} \times 0.05\right) \text{ g}$$

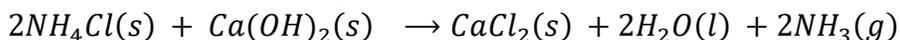
$$= 3.2 \text{ g of copper}$$

**∴ The mass of copper formed is 3.2 g**

3. Calculate the mass of ammonium chloride that will just react completely with 14.8g of calcium hydroxide. ( $N = 14; H = 1; Cl = 35.5$ )

$$\text{Molar mass of } \text{NH}_4\text{Cl} = (14 \times 1) + (1 \times 4) + (35.5 \times 1) = 53.5\text{g}$$

$$\text{Molar mass of } \text{Ca(OH)}_2 = (40 \times 1) + (16 \times 2) + (1 \times 2) = 74\text{g}$$



1 mole of  $\text{Ca(OH)}_2$  reacts with 2 moles of  $\text{NH}_4\text{Cl}$

74 g of  $\text{Ca(OH)}_2$  react with  $(2 \times 53.5)\text{g}$  of  $\text{NH}_4\text{Cl}$

$$14.8 \text{ g of } \text{Ca(OH)}_2 \text{ react with } \left(\frac{2 \times 53.5}{74} \times 14.8\right) \text{ g of } \text{NH}_4\text{Cl}$$

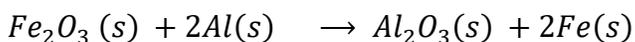
$$= 21.4 \text{ g of } \text{NH}_4\text{Cl}$$

**∴ The mass of Ammonium chloride that will react = 21.4 g**

4. Calculate the mass of copper(II) hydroxide which could be formed by reacting 72g of copper(II) sulphate with excess sodium hydroxide solution.

$$(\text{Cu} = 64; \text{S} = 32; \text{Na} = 23; \text{O} = 16; \text{H} = 1)$$

5. Aluminium reacts with iron(III) oxide according to the following equation.



In the above reaction, 10g of iron were formed. Calculate the;

- a) The number of moles of iron(III) oxide that reacted with aluminium.  
( $Fe = 56; Al = 27; O = 16$ )
- b) The mass of iron(III) oxide that reacted.
6. Determine the mass of iodine in 5g of lead(II) iodide. ( $Pb = 207; I = 127$ )

***Calculating Volumes from stoichiometric equations***

1. 50g of calcium carbonate completely reacted with hydrochloric acid. Calculate the volume of carbon dioxide liberated at s.t.p.

( $Ca = 40; C = 12; O = 16; H = 1$ ; 1 mole of a gas occupies  $22.4dm^3$ )

***Method 1***

*Moles of calcium carbonate that reacted*

Molar mass of  $CaCO_3 = (40 \times 1) + (12 \times 1) + (16 \times 3) = 100g$

100g of calcium carbonate contain 1 mole.

50 g of calcium carbonate contain  $\left(\frac{1}{100} \times 50\right)$  moles

= 0.5 moles of calcium carbonate.

*Moles of carbon dioxide liberated*



From the equation;

1 mole of calcium carbonate liberates 1 mole of carbon dioxide.

0.5 moles of calcium carbonate liberate  $\left(\frac{1}{1} \times 0.5\right)$  moles of carbon dioxide

= 0.5 moles of carbon dioxide.

*Volume of carbon dioxide liberated.*

1 mole of carbon dioxide occupies  $22.4dm^3$  at s. t. p.

0.5 moles of carbon dioxide occupy  $\left(\frac{22.4}{1} \times 0.5\right)dm^3$

=  $11.2 dm^3$  of carbon dioxide.

***∴ The volume of carbon dioxide liberated =  $11.2dm^3$***

***Method 2***

Molar mass of  $CaCO_3 = (40 \times 1) + (12 \times 1) + (16 \times 3) = 100g$



From the equation;

1 mole of calcium carbonate liberated 1 mole of carbon dioxide.

100g of calcium carbonate liberated 22.4dm<sup>3</sup> of carbon dioxide.

50g of calcium carbonate will liberate  $\left(\frac{22.4}{100} \times 50\right)$  dm<sup>3</sup> of carbon dioxide  
 = 11.2dm<sup>3</sup> of carbon dioxide

**∴ The volume of carbon dioxide liberated = 11.2dm<sup>3</sup>**

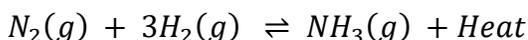
2. Copper(II) carbonate reacts with dilute hydrochloric acid according to the equation;



Calculate the volume of carbon dioxide that would be evolved at s.t.p when 6.2g of copper(II) carbonate is reacted with excess dilute hydrochloric acid.

(H = 1; C = 12; Cu = 64; 1 mole of gas occupies 22.4dm<sup>3</sup> at s. t. p)

3. Nitrogen reacts with hydrogen to produce ammonia according to the following equation;



Calculate the volume of ammonia produced at s.t.p when 18.5g of hydrogen gas is reacted with excess nitrogen.

(H = 1; N = 14; 1 mole of gas occupies 22.4 litres at s. t. p)

4. Hydrogen chloride reacts with lead(II) nitrate to form lead(II) chloride according to the following equation;



Calculate the volume of hydrogen chloride measured at s.t.p that would be required to form 5.53g of lead(II) chloride.

(PbCl<sub>2</sub> = 277; 1 mole of gas occupies 22.4dm<sup>3</sup> at s. t. p)

**(Answer; volume = 0.8944dm<sup>3</sup>)**

5. Calculate the volume of the total gaseous product formed at room temperature when 4.5g of calcium nitrate is heated strongly.

(N = 14; O = 16; Ca = 40; 1 mole of gas occupies 24.0dm<sup>3</sup> at room temperature)

**(Answer; volume = 1.6464dm<sup>3</sup>)**

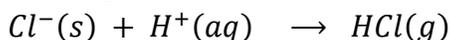
6. (a) Soot is a form of carbon.

i. Write the equation for the complete combustion of soot in oxygen.

- ii. Calculate the volume of the gas produced at room temperature when 0.6g of soot is burnt in excess oxygen.  
 ( $C = 12$ ; 1 mole of gas occupies  $24.0\text{dm}^3$  at room temperature)  
**(Answer; Volume =  $1.2\text{dm}^3$ )**

***Calculating mass using volumes and stoichiometric equations***

1. Hydrogen chloride can be prepared from sodium chloride according to the following ionic equation;



Calculate the mass of sodium chloride that would be required to produce  $3.60\text{dm}^3$  of hydrogen chloride measured at room temperature.

( $H = 1$ ;  $Na = 23$ ;  $Cl = 35.5$ ; one mole of gas occupies  $24.0\text{dm}^3$ )

Moles of hydrogen chloride produced

$24.0\text{dm}^3$  are occupied by 1 mole of hydrogen chloride

$3.60\text{dm}^3$  will be occupied by  $\left(\frac{1}{24.0} \times 3.60\right)$  moles of hydrogen chloride.  
**= 0.15 moles of hydrogen chloride.**

Moles sodium chloride that reacted

From the equation above;

1 mole of hydrogen chloride was produced by 1 mole of sodium chloride.

0.15 moles of hydrogen chloride produced by  $\left(\frac{1}{1} \times 0.15\right)$  moles of NaCl  
**= 0.15 moles of sodium chloride.**

Mass of sodium chloride used

Molar mass of NaCl =  $(23 \times 1) + (35.5 \times 1) = 58.5\text{ g}$

1 mole of sodium chloride weighs  $58.5\text{g}$

**0.15 moles of sodium chloride will weigh  $\left(\frac{58.5}{1} \times 0.15\right)\text{g}$**   
 =  $8.775\text{g}$  of sodium chloride  
 $\approx 8.8\text{ g}$  of sodium chloride.

**$\therefore$  The mass of sodium chloride that would be required =  $8.8\text{ g}$**

**ALTERNATIVELY;**

Molar mass of NaCl =  $(23 \times 1) + (35.5 \times 1) = 58.5\text{ g}$

From the equation above;

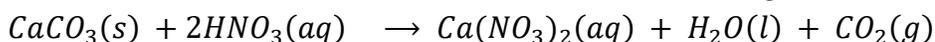
1 mole of hydrogen chloride was produced by 1 mole of sodium chloride.

24.0dm<sup>3</sup> of hydrogen chloride was produced by 58.5 g of sodium chloride

3.60dm<sup>3</sup> of hydrogen chloride would be produced by  $\left(\frac{58.5}{24.0} \times 3.60\right)$  g of NaCl  
= 8.775g of sodium chloride.  
≈ 8.8g of sodium chloride

∴ **The mass of sodium chloride that would be required = 8.8 g**

2. Calcium carbonate reacts with dilute nitric acid according to the following equation;



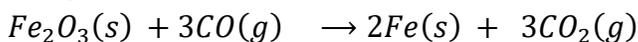
Calculate the mass of calcium carbonate that would liberate 250cm<sup>3</sup> of carbon dioxide at room temperature.

(Ca = 40; C = 12; O = 16; 1 mole of gas occupies 24dm<sup>3</sup> at room temperature)

**(Answer; mass = 1.04g)**

### **Activity**

1. 4.20g of heated iron(III) oxide was treated with carbon monoxide as follows;

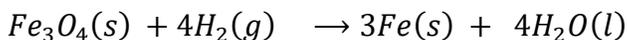


Calculate the volume of carbon dioxide evolved at room temperature.

(Fe = 56; O = 16; C = 12; 1 mole of gas occupies 24dm<sup>3</sup> at room temperature)

**(Answer; volume = 1.89dm<sup>3</sup>)**

2. Hydrogen reacts with iron(II,III) oxide according to the following equation.

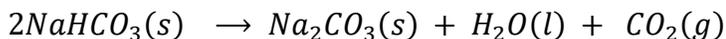


Calculate the volume of hydrogen measured at room temperature that would be required to produce 3.36g of iron

(Fe = 56; O = 16; H = 1; 1 mole of gas occupies 24dm<sup>3</sup> at room temperature)

**(Answer; volume = 1.92dm<sup>3</sup>)**

3. Carbon dioxide can be prepared from sodium hydrogen carbonate according to the equation;

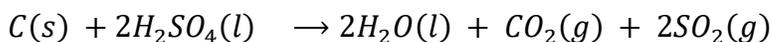


Calculate the volume in litres of carbon dioxide evolved at s.t.p when 21.0g of sodium hydrogen carbonate is heated.

(Na = 23; H = 1; C = 12; O = 16; molar gas volume at s. t. p = 22.4l)

**(Answer; volume = 2.8l)**

4. Concentrated Sulphuric acid reacts with graphite according to the following equation;



Calculate the mass of carbon that can react completely with a solution containing 19.6g of Sulphuric acid. ( $C = 12; S = 32; H = 1; O = 16$ )

**(Answer; mass = 1.2g)**

5. When 3.0g of X was heated, 210cm<sup>3</sup> of a gas were evolved at s.t.p and 2.4g of the solid residue remained. Calculate the relative molecular mass of the gas.

(1 mole of gas occupies 22.4dm<sup>3</sup> at s. t. p)

**(Answer; RMM of gas = 64)**

6. Sodium hydrogen carbonate was heated strongly.

a) Write the equation for the reaction that occurred.

b) Calculate the mass of sodium hydrogen carbonate that must be heated to give of 200cm<sup>3</sup> of carbon dioxide at room temperature.

(1 mole of gas occupies 24000cm<sup>3</sup> at r. t.;  $Na = 23; O = 16; C = 12; H = 1$ )

**(Answer; mass = 1.4g)**

7. Hydrogen peroxide decomposes when exposed to sun light.

a) Write the equation for the reaction that occurred.

b) Determine the volume of oxygen formed when 24.8g of hydrogen peroxide is completely decomposed at s.t.p

( $H = 1; O = 16$ ; 1 mole of gas occupies 22.4dm<sup>3</sup> at s. t. p)

**(Answer; volume = 8.2dm<sup>3</sup>)**

8. Calculate the volume of oxygen at s.t.p which could be obtained by heating 5g of potassium chlorate. ( $K = 39; Cl = 35.5; O = 16$ ; Molar gas volume at s. t. p = 22.4l)

**(Answer; volume = 1.37l)**

9. What volume of ammonia gas will be oxidized by 60g of copper(II) oxide at s.t.p?

( $Cu = 64; O = 16; H = 1; N = 14$ ; molar gas volume at s. t. p = 22400cm<sup>3</sup>)

10. What volume of carbon dioxide at s.t.p would be obtained by dissolving 150g of pure calcium carbonate in dilute nitric acid?

( $Ca = 40; O = 16; C = 12$ ; ; molar gas volume at s. t. p = 22.4dm<sup>3</sup>)

11. Calculate the volume of ammonia at s.t.p that would be produced when 5g of hydrogen react with nitrogen.

( $H = 1; N = 14$ ; molar gas volume at s. t. p = 22400cm<sup>3</sup>)

12. When excess dilute sulphuric acid was added to 4.50g of a mixture containing sodium sulphate and sodium sulphite, 720cm<sup>3</sup> of sulphur dioxide was collected at room temperature. Calculate the percentage of sodium sulphate that was in the mixture.  
( $Na_2SO_4 = 142$ ;  $Na_2SO_3 = 126$ ; 1 mole of gas occupies 24dm<sup>3</sup> at r. t)  
(Answer; %age = 16%)
13. A stream of excess dry hydrogen was passed over 6.85g of heated lead(II) oxide in a combustion tube. ( $Pb = 207$ ;  $O = 16$ )
- State what was observed.
  - Write the equation for the reaction.
  - Determine the mass of the residue formed.
  - Work out the number of moles of water vapour formed.
14. A solid nitrate Q consists of 9.37% magnesium, 10.93% nitrogen, 42.18% of water and the rest being oxygen. Determine the simplest formula of Q ( $Mg = 24$ ;  $N = 14$ ;  $O = 16$ ;  $H = 1$ )
15. A gaseous hydrocarbon N consists of 82.7% carbon by mass. 0.36g of N occupies 139cm<sup>3</sup> at s.t.p. Determine the molecular formula of N  
( $C = 12$ ;  $H = 1$ ; 1 mole of gas occupies 22400cm<sup>3</sup> at s. t. p)
16. Lead(II) nitrate was heated strongly.
- Write the equation for the reaction that took place.
  - Calculate the mass of lead(II) nitrate to be heated to form 1.5dm<sup>3</sup> of nitrogen dioxide gas at s.t.p.  
(RFM of  $Pb(NO_3)_2 = 331$ ; molar gas volume at s. t. p = 22.4dm<sup>3</sup>)  
(Answer; mass = 11.09g)
17. A sample of potassium nitrate crystals was heated strongly until no further change
- Write the equation for the reaction that took place.
  - Calculate the loss in mass if 20.2g of potassium nitrate crystals was completely decomposed. ( $K = 39$ ;  $N = 14$ ;  $O = 16$ )
18. Magnetite,  $Fe_3O_4$  can be converted to iron using carbon monoxide and hydrogen separately;
- Write equation for the reaction between heated magnetite and;
    - Dry hydrogen gas.
    - Dry carbon monoxide

## GAY LUSSAC'S LAW

State 'Gay Lussac's law'

- *The law states that **when gases react, they do so in volumes which bear a simple ratio to one another and to the volume of the gaseous products, provided all the volumes are measured at the same temperature and pressure.***

*Note; The law does not apply to solids or liquids.*

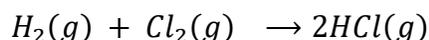
Gay Lussac's studied the chemical reactions between gases and noticed that there is always a very simple ratio between the volumes of gases that react together. For instance;

- $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$   
2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of steam.
- $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$   
1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrogen chloride gas.
- $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$   
1 volume of nitrogen combines with 3 volume of hydrogen to form 2 volumes of ammonia gas.

### Calculations involving use of Gay Lussac's law

1. 100cm<sup>3</sup> of hydrogen were burnt in excess chlorine. Determine;

a) the volume of chlorine used.



1 volume of hydrogen combines with 1 volume of chlorine

$$100\text{cm}^3 \text{ of hydrogen combine with } \left(\frac{1}{1} \times 100\right) \text{cm}^3 \text{ of chlorine} \\ = 100\text{cm}^3 \text{ of chlorine}$$

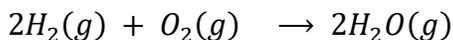
b) The volume of the gaseous product formed.

From the equation;

1 volume of hydrogen produces 2 volumes of hydrogen chloride gas.

$$100\text{cm}^3 \text{ of hydrogen produce } \left(\frac{2}{1} \times 100\right) \text{cm}^3 \text{ of hydrogen chloride gas} \\ = 200\text{cm}^3 \text{ of hydrogen chloride gas.}$$

2. 80cm<sup>3</sup> of oxygen are exploded with 80cm<sup>3</sup> of hydrogen. Determine the volume of the unused (residual gas) at room temperature.



2 volumes of hydrogen combine with 1 volume of oxygen .

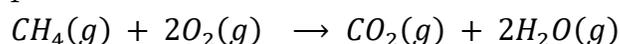
$80\text{cm}^3$  of hydrogen combine with  $\left(\frac{1}{2} \times 80\right)\text{cm}^3$  of oxygen  
 $= 40\text{cm}^3$  of oxygen.

volume of oxygen that reacted  $= 40\text{cm}^3$

Volume of unused gas(oxygen) = original volume – volume that reacted.  
 $= 80 - 40$   
 $= 40\text{cm}^3$  of oxygen.

$\therefore$  **Volume of unused gas =  $40\text{cm}^3$**

3.  $48\text{cm}^3$  of methane was exploded with  $212\text{cm}^3$  of oxygen at a certain temperature and pressure. Determine the volume of the remaining gases.



1 volume of methane combines with 2 volumes of oxygen.

$48\text{cm}^3$  of methane combine with  $\left(\frac{2}{1} \times 48\right)\text{cm}^3$  of oxygen.  
 $= 96\text{cm}^3$  of oxygen.

Volume of un reacted oxygen  $= 212 - 96$   
 $= 116\text{cm}^3$

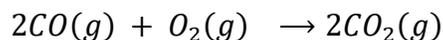
1 volume of methane produced 1 volume of carbon dioxide.

$48\text{cm}^3$  of methane produce  $\left(\frac{1}{1} \times 48\right)\text{cm}^3$  of carbon dioxide.  
 $= 48\text{cm}^3$  of carbon dioxide

Volume of remaining gases = volume of un reacted oxygen + volume of  $\text{CO}_2$   
 $= 116 + 48$   
 $= 164\text{cm}^3$

**N.B. At room temperature, the 2 volumes of steam produced condensed to water.**

4.  $10\text{cm}^3$  of carbon monoxide are mixed with  $20\text{cm}^3$  of oxygen, ignited and allowed to cool to the original temperature and pressure. Calculate the total volume of the gases after the reaction.



2 volumes of carbon monoxide combine with 1 volume of oxygen.

$10\text{cm}^3$  of carbon monoxide combine with  $\left(\frac{1}{2} \times 10\right)\text{cm}^3$  of oxygen.  
 $= 5\text{cm}^3$  of oxygen.

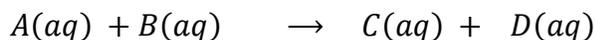
## **VOLUMETRIC ANALYSIS**

In an experiment to investigate or determine the nature of matter in a substance, a chemist focuses on two main questions;

- a) What are the components of the substance?
- b) What amount of each component is present in a substance?

In an attempt to answer question (a) and (b), a chemist carries out qualitative and quantitative analysis respectively. Most of the reactions which a chemist carries out take place in solution.

Consider the reaction below;



If we determine the volume of A needed to completely react with a given amount or volume of B, the answer is provided practically through volumetric analysis.

In volumetric analysis, quantities of substances (often acids or alkalis) are estimated by analytical process involving measurements of volumes of solutions using pipettes, burettes and measuring cylinders (for approximate measurements). Weighing may be also involved. Most of the work in volumetric analysis is based upon molar(*M*) Solutions.

### ***STANDARD AND MOLAR SOLUTIONS***

A **standard solution** is a solution of known concentration. Examples of standard solutions are;

- Solution containing 12g of sodium chloride in one litre of solution.
- A solution containing 2 moles of sodium hydroxide in 1dm<sup>3</sup>. etc.

*Note*; The substance that is used to prepare a standard solution is known as a **primary standard**.

A **molar solution** is a solution that contains one mole of a substance in a solution of one litre.

In other words, **molar solution** is a solution containing one mole of solute in one litre.

***Other related terms are;***

**Concentration**; This is the amount of substances/solutes in a given volume of solution.

**Molarity**; This is the number of moles of substance/ solute in one litre of solution.

Or Molarity is the concentration of a substance in moles per litre.

### ***UNITS OF MOLARITY***

Moles per litre or Mol/l or  $M\text{ol}^{-1}$

Moles per cubic decimetre or  $\text{Mol}/\text{dm}^3$  or  $M\text{oldm}^{-3}$

Molarity of a solution is denoted by letter  $M$  E.g.

0.2M sodium hydroxide means that 1 litre of solution containing 0.2 moles of NaOH.

1 litre = 1 cubic decimetre = 1000 cubic centimetre.

### **CALCULATIONS ON MOLARITY AND MASSES**

#### **Examples.**

1. 2g of sodium hydroxide were dissolved in water to make 200 $\text{cm}^3$ . Calculate the molarity of the solution formed. ( $Na = 23$ ;  $O = 16$ ;  $H = 1$ )

#### **Method 1**

$$\begin{aligned}\text{Molar mass of NaOH} &= (23 \times 1) + (16 \times 1) + (1 \times 1) \\ &= 23 + 16 + 1 \\ &= 40\text{gmol}^{-1}\end{aligned}$$

40g of Sodium hydroxide contain 1 mole

$$\begin{aligned}2\text{g of Sodium hydroxide contain } &\left(\frac{1}{40} \times 2\right) \text{ moles} \\ &= 0.05 \text{ moles of sodium hydroxide.}\end{aligned}$$

200 $\text{cm}^3$  of solution contain 0.05 moles of Sodium hydroxide.

$$\begin{aligned}1000\text{cm}^3 \text{ of solution contain } &\left(\frac{0.05}{200} \times 1000\right) \text{ moles of sodium hydroxide.} \\ &= 0.25 \text{ moles of sodium hydroxide.}\end{aligned}$$

**$\therefore$  The molarity of solution is 0.25M**

#### **Method 2**

200 $\text{cm}^3$  of solution contain 2g of Sodium hydroxide

$$\begin{aligned}1000\text{cm}^3 \text{ of solution contain } &\left(\frac{2}{200} \times 1000\right) \text{ g of sodium hydroxide} \\ &= 10\text{gl}^{-1} \text{ of sodium hydroxide.}\end{aligned}$$

$$\begin{aligned}\text{Molar mass of NaOH} &= (23 \times 1) + (16 \times 1) + (1 \times 1) \\ &= 23 + 16 + 1 \\ &= 40\text{gmol}^{-1}\end{aligned}$$

40g of Sodium hydroxide contain 1 mole

$$\begin{aligned}10\text{g of Sodium hydroxide contain } &\left(\frac{1}{40} \times 10\right) \text{ moles} \\ &= 0.25 \text{ moles of sodium hydroxide}\end{aligned}$$

**∴ The molarity of solution is 0.25M**

2. A solution of calcium iodide was made by dissolving 29.4g in one litre of solution. Determine the molarity of the solution. ( $Ca = 40; I = 127$ )

$$\begin{aligned} \text{Molar mass of } CaI_2 &= (40 \times 1) + (127 \times 2) \\ &= 40 + 254 \\ &= 294 \text{ gmol}^{-1} \end{aligned}$$

294g of calcium iodide contain 1 mole

$$\begin{aligned} 29.4 \text{ g of calcium iodide contain } &\left(\frac{1}{294} \times 29.4\right) \text{ moles} \\ &= 0.1 \text{ mole of calcium iodide} \end{aligned}$$

**∴ The molarity of the solution is 0.1M**

3. 5.72g of hydrated sodium carbonate ( $Na_2CO_3 \cdot 10H_2O$ ) were dissolved in water to make 500cm<sup>3</sup> of solution. Determine the molarity of the solution.

( $Na = 23; C = 12; O = 16; H = 1$ )

$$\begin{aligned} \text{Molar mass of } Na_2CO_3 \cdot 10H_2O &= (23 \times 2) + (12 \times 1) + (16 \times 3) + (1 \times 2 \times 10) + (16 \times 1 \times 10) \\ &= 286 \text{ gmol}^{-1} \end{aligned}$$

286g of sodium carbonate – 10 – water contain 1 mole

$$\begin{aligned} 5.72 \text{ g of sodium carbonate – 10 – water contain } &\left(\frac{1}{286} \times 5.72\right) \text{ moles} \\ &= 0.02 \text{ moles} \end{aligned}$$

500cm<sup>3</sup> of solution contain 0.02 moles of sodium carbonate – 10 – water

$$\begin{aligned} 1000 \text{ cm}^3 \text{ of solution contain } &\left(\frac{0.02}{500} \times 1000\right) \text{ moles of } Na_2CO_3 \cdot 10H_2O \\ &= 0.04 \text{ moles of } Na_2CO_3 \cdot 10H_2O \end{aligned}$$

**∴ The molarity of the solution is 0.04M**

4. 2.25g of anhydrous oxalic acid was dissolved in water to make 400cm<sup>3</sup> of solution. Determine the concentration of oxalic acid in moles per litre.

$$\begin{aligned} \text{Molar mass of } H_2C_2O_4 &= (1 \times 2) + (12 \times 2) + (16 \times 4) \\ &= 90 \text{ gmol}^{-1} \end{aligned}$$

90g of oxalic acid contain 1 mole.

$$\begin{aligned} 2.25 \text{ g of oxalic acid contain } &\left(\frac{1}{90} \times 2.25\right) \text{ moles} \\ &= 0.025 \text{ moles of oxalic acid.} \end{aligned}$$

400cm<sup>3</sup> of solution contain 0.025 moles of oxalic acid.

$$\begin{aligned} 1000 \text{ cm}^3 \text{ of solution contain } &\left(\frac{0.025}{400} \times 1000\right) \text{ moles of oxalic acid.} \\ &= 0.0625 \text{ moles of oxalic acid.} \end{aligned}$$

**∴ The concentration of oxalic acid is 0.0625 moles per litre.**

5. The molarity of potassium hydroxide is 0.01M. Calculate the concentration of potassium hydroxide in grams per litre. ( $K = 39; O = 16; H = 1$ )

$$\begin{aligned} \text{Molar mass of KOH} &= (39 \times 1) + (16 \times 1) + (1 \times 1) \\ &= 56 \text{ g mol}^{-1} \end{aligned}$$

1 mole of potassium hydroxide contain 56g

$$\begin{aligned} 0.01 \text{ mole of potassium hydroxide contain } &\left(\frac{56}{1} \times 0.01\right) \text{ g} \\ &= 0.56 \text{ g of potassium hydroxide.} \end{aligned}$$

**$\therefore$  The concentration of potassium hydroxide is 0.56g l<sup>-1</sup>**

6. The molarity of a solution X is 0.05M. Calculate the number of moles contained in 25cm<sup>3</sup> of X.

1000cm<sup>3</sup> of solution contain 0.05 moles of X

$$\begin{aligned} 25 \text{ cm}^3 \text{ of solution contain } &\left(\frac{0.05}{1000} \times 25\right) \text{ moles of X} \\ &= 0.00125 \text{ moles of X.} \end{aligned}$$

7. The concentration of potassium hydroxide solution is 1.12g/l. Calculate the number of moles of potassium hydroxide contained in 20cm<sup>3</sup> of that solution.

( $K = 39; H = 1; O = 16$ )

$$\begin{aligned} \text{Molar mass of KOH} &= (39 \times 1) + (16 \times 1) + (1 \times 1) \\ &= 56 \text{ g mol}^{-1} \end{aligned}$$

56g of potassium hydroxide contain 1 mole

$$\begin{aligned} 1.12 \text{ g of potassium hydroxide contain } &\left(\frac{1}{56} \times 1.12\right) \text{ moles} \\ &= 0.02 \text{ moles of KOH per litre.} \end{aligned}$$

1000cm<sup>3</sup> of solution contain 0.02 moles of potassium hydroxide.

$$\begin{aligned} 20 \text{ cm}^3 \text{ of solution contain } &\left(\frac{0.02}{1000} \times 20\right) \text{ moles of potassium hydroxide.} \\ &= 0.0004 \text{ moles of potassium hydroxide.} \end{aligned}$$

**Alternatively;**

1000cm<sup>3</sup> of solution contain 1.12g of potassium hydroxide.

$$\begin{aligned} 20 \text{ cm}^3 \text{ of solution contain } &\left(\frac{1.12}{1000} \times 20\right) \text{ g of potassium hydroxide.} \\ &= 0.0224 \text{ g of potassium hydroxide.} \end{aligned}$$

$$\begin{aligned} \text{Molar mass of KOH} &= (39 \times 1) + (16 \times 1) + (1 \times 1) \\ &= 56 \text{ g mol}^{-1} \end{aligned}$$

56g of potassium hydroxide contain 1 mole

$$\begin{aligned} 0.0224 \text{ g of potassium hydroxide contain } &\left(\frac{1}{56} \times 0.0224\right) \text{ moles} \\ &= 0.0004 \text{ moles of KOH} \end{aligned}$$

8. Determine the mass of sulphuric acid in  $5\text{cm}^3$  of  $0.2\text{M}$  solution of the acid.  
( $H = 1; S = 32; O = 16$ )

$1000\text{cm}^3$  of solution contain  $0.2$  moles of sulphuric acid.

$5\text{cm}^3$  of solution contain  $\left(\frac{0.2}{1000} \times 5\right)$  moles of sulphuric acid.  
 $= 0.001$  moles of sulphuric acid.

Molar mass of  $\text{H}_2\text{SO}_4 = (1 \times 2) + (32 \times 1) + (16 \times 4)$   
 $= 98\text{g mol}^{-1}$

1 mole of sulphuric acid weighs  $98\text{g}$

$0.001$  mole of sulphuric acid weigh  $\left(\frac{98}{1} \times 0.001\right)\text{g}$   
 $= 0.098\text{g}$  of sulphuric acid.

**In general, Molarity =  $\frac{\text{Concentration(g/l)}}{\text{Molar mass}}$**

When the concentration of a solution in grams per litre and the RMM are known, then the molarity can be calculated from the above expression.

**N. B. The use of the formula is not so much recommended and workings should be from first principles.**

#### Activity

1. Calculate the concentration in moles per litre of solution containing  $0.4$  moles of sodium hydroxide dissolved in water and the solution made up to  $100\text{cm}^3$ .
2. Calculate the mass of oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ , in grams required to prepare  $250\text{cm}^3$  of  $1.5\text{M}$  solution of oxalic acid. ( $H = 1; C = 12; O = 16$ )  
(Answer;  $33.75\text{g}$ )
3. Calculate the concentration in grams per litre of a solution containing  $0.05$  moles of sodium hydroxide in  $50\text{cm}^3$ . ( $\text{Na} = 23; \text{Cl} = 35.5$ )  
(Answer;  $58.5\text{g/l}$ )
4.  $0.4\text{g}$  of a metal hydroxide,  $\text{MOH}$  reacted completely with  $20\text{cm}^3$  of  $0.5\text{M}$  hydrochloric acid. Determine the relative atomic formula mass of  $\text{MOH}$  hence deduce the atomic mass of  $\text{M}$ . ( $O = 16; H = 1$ )
5. Determine the concentration of chloride ions in a litre of a solution containing  $22.2\text{g}$  of calcium chloride. ( $\text{Ca} = 40; \text{Cl} = 35.5$ )

(Answer;  $0.40\text{mol l}^{-1}$ )

## PREPARATION OF SOLUTIONS

**Standard solution:** Is a solution whose concentration is precisely known.

A standard solution is prepared from a substance which can be obtained in its pure form called a primary standard.

**Primary standard:** Is a chemical substance that is analytically pure and chemically stable in aqueous solution which when a known mass of it is dissolved in known volume of solvent, a standard solution is prepared and can be used to standardize other solutions.

**Or**

A Primary standard is a substance which is pure and stable, used to prepare a standard solution by dissolving a known mass of solid in a certain quantity of solvent.

Its stability in this case is used to mean that the concentration of the substance does not change with time.

### Characteristics/properties of primary standard

- ❖ It should be readily available in state of highest purity.
- ❖ It should have high relative molecular mass to reduce weighing errors.
- ❖ It should be readily soluble in water.
- ❖ It should be stable in storage and on exposure to air I.e. it should not be hygroscopic, efflorescence, affected by oxygen or carbon dioxide.

### Examples of primary standards

- ❖ Anhydrous sodium carbonate.
- ❖ Disodium tetra borate decahydrate (sodium tetra borate)  $Na_2B_4O_7 \cdot 10H_2O$ .
- ❖ Ethane dioic acid.
- ❖ Butane dioic acid

- ❖ Oxalic acid.
- ❖ Benzoic acid.
- ❖ Constant boiling hydrochloric acid.
- ❖ Sodium hydrogen carbonate.
- ❖ Potassium dichromate(VI).
- ❖ Sodium Oxalate.
- ❖ Ferrous ammonium sulphate.
- ❖ Potassium Iodate.
- ❖ Iodine
- ❖ Potassium or sodium ethanoate.etc.

### ***Secondary standard***

Any substance/ compound which does not have the above properties of primary standard is referred to as a **secondary standard**. Solutions of secondary standards are used in volumetric analysis but their results obtained are not normally accurate.

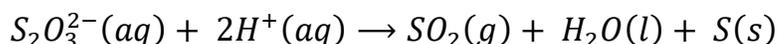
The following compounds are **NOT** primary standards;

#### **1. Sodium hydroxide;**

- ❖ Because it is deliquescent (absorbs water from the atmosphere) and dissolves in it, forming a solution.
- ❖ The solution absorbs carbon dioxide forming a solution of sodium carbonate.
- ❖ In excess carbon dioxide white powder of sodium hydrogen carbonate is formed.

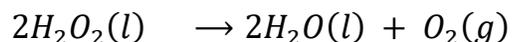
#### **2. Sodium thiosulphate-5-water.**

- ❖ Has variable amount of water of crystallization.
- ❖ When exposed to air its solution absorbs carbon dioxide to form a weak carbonic acid which ionizes to give hydrogen ions, the hydrogen ions then react with thiosulphate ions to form sulphur, sulphur dioxide and water.



#### **3. Hydrogen peroxide.**

- ❖ Because it is unstable at room temperature as it easily decomposes to give oxygen and water.

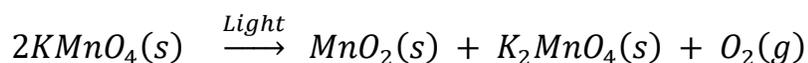


4. **Hydrated sodium carbonate.** ( $Na_2CO_3 \cdot 10H_2O$ )

- ❖ Because it is efflorescence. (it loses its water to the atmosphere)

5. **Potassium manganate(VII)**

- ❖ Because it is difficult to obtain it in the pure state, and it is slowly reduced by water to manganese(IV) oxide especially in the presence of light or acid.



- ❖ Its solution slowly deposits brown manganese(IV) oxide on standing.

Concentrated hydrochloric acid, concentrated nitric acid and concentrated ammonia are unsuitable because they are volatile

Concentrated Sulphuric acid is also unsuitable because it is hygroscopic i.e. absorbs moisture from the atmosphere.

Calcium hydroxide and Magnesium hydroxide are unsuitable as primary standards because they are less soluble in water.

***PREPARATION OF STANDARD SOLUTIONS.***

Standard solutions may be prepared by either weighing a known mass, dissolving it and diluting to a given volume or by diluting a more concentrated solution of known concentration.

**(a) Preparation of a standard solution by weighing.**

Example;

Preparing 250cm<sup>3</sup> of 0.25M sodium hydrogen carbonate solution.

*Procedure*

- ✓ First calculate the mass of sodium hydrogen carbonate needed.

$$\begin{aligned} \text{Molar mass of NaHCO}_3 &= 23 + 1 + 12 + (16 \times 3) \\ &= 84g \end{aligned}$$

1 mole of  $\text{NaHCO}_3$  weighs 84 g

$$0.25 \text{ mole of } \text{NaHCO}_3 \text{ weigh } \left(\frac{84}{1} \times 0.25\right) \text{ g}$$
$$= 21 \text{ g of } \text{NaHCO}_3$$

1000cm<sup>3</sup> of solution contain 21 g of  $\text{NaHCO}_3$

$$250\text{cm}^3 \text{ of solution will contain } \left(\frac{21}{1000} \times 250\right) \text{ g of } \text{NaHCO}_3$$
$$= 5.25 \text{ g of } \text{NaHCO}_3$$

- ✓ Weigh accurately 5.25g of  $\text{NaHCO}_3$  into a clean beaker.
- ✓ Add about 100cm<sup>3</sup> of distilled water. Stir with a glass rod until all the solid has dissolved.
- ✓ Transfer all the solution carefully through a filter funnel into a 250cm<sup>3</sup> volumetric flask. Wash all the solution out of the beaker and off the glass rod.
- ✓ Add distilled water until the level is about 2cm<sup>3</sup> below the graduation mark on the volumetric flask. Add the rest of the distilled water drop by drop from a dropping pipette or wash bottle until the bottom of the meniscus is level with the graduation mark when viewed at eye level. Insert the stopper of the flask and invert the flask several times to mix the solution. The solution can be labeled say FA1. The results are recorded in the table as below.

Mass of beaker + solid Q= 38.75g

Mass of beaker alone =33.55g

Mass of solid Q =5.25g

#### EXERCISE:

1. Calculate (to 2 decimal places) the mass of sodium carbonate needed to be weighed to prepare 500ml of 0.15M sodium carbonate solution. **(Ans.7.95g)**

- Calculate the mass of sodium hydrogen carbonate that would be required to prepare  $250\text{cm}^3$  of  $0.2\text{M}$  sodium hydrogen carbonate solution.
- $2\text{g}$  of sodium hydrogen carbonate were dissolved in  $400\text{cm}^3$  of water. What is the molarity of the solution?

***Diluting a stock solution.***

Concentrated solutions from which the diluted solutions can be prepared are called stock solutions.

Dilution refers to decreasing the concentration of a substance by increasing the volume of the solvent.

Stock solutions of sulphuric acid, nitric acid, hydrochloric acid, ammonia e.t.c are sold in high concentrations and kept in well labeled containers. They are normally used in laboratory to prepare dilute solutions of lower concentrations.

The concentrations of stock solutions are normally expressed as percentage weight by volume ( $W/V$ ) or percentage weight by weight ( $W/W$ ) with the following data on the container; percentage purity, density and formula mass. Concentrated nitric acid is  $69\%$  ( $W/W$ ), it means that there are  $69\text{g}$  of nitric acid in  $100\text{g}$  of acid.

Examples:

- A bottle containing stock solution of concentrated sulphuric acid has the following data;

Stock bottle	Density ( $\text{gcm}^{-3}$ )	Formula mass(g)	% purity
$H_2SO_4$	1.835	98	98

Determine the molarity of the stock solution (concentrated sulphuric acid.)

$1\text{cm}^3$  of solution contains  $1.835\text{g}$  of  $H_2SO_4$

$$1000\text{cm}^3 \text{ of solution contain } \left( \frac{1.835}{1} \times 1000 \right) \text{ g of } H_2SO_4$$

$$= 1835\text{gl}^{-1}$$

$$\begin{aligned} \text{Mass of pure } H_2SO_4 &= \left(\frac{98}{100} \times 1835\right) g \\ &= 1798.3 g \text{ of } H_2SO_4 \end{aligned}$$

98g of pure  $H_2SO_4$  in  $1000\text{cm}^3$  contain 1.0 mole.

$$\begin{aligned} 1798.3 g \text{ of pure } H_2SO_4 \text{ in } 1000\text{cm}^3 \text{ contain } &\left(\frac{1.0}{98} \times 1798.3\right) \text{ moles} \\ &= 18.35 \text{ moles} \end{aligned}$$

**$\therefore$  Molarity of 98% conc.  $H_2SO_4 = 18.35M$**

N.B: The volume of the stock solution that should be taken to dilute to the desired concentration is calculated. i.e.

**Preparing one litre of 0.1M sulphuric acid**

18.35 moles of  $H_2SO_4$  solution are contained in  $1000\text{cm}^3$

$$\begin{aligned} 0.1 \text{ mole of } H_2SO_4 \text{ solution are contained in } &\left(\frac{1000}{18.35} \times 0.1\right) \text{ cm}^3 \\ &= 5.44\text{cm}^3 \end{aligned}$$

$\therefore 5.44\text{cm}^3$  of the stock solution is measured and placed in a  $1000\text{cm}^3$  volumetric flask containing about  $500\text{cm}^3$  of distilled water. Then distilled water is added to the mixture up to the mark. Volumetric flask shaken to obtain a uniform solution.

This solution is 0.1M sulphuric acid.

**EXERCISE**

1. Concentrated nitric acid is 70%w/W and has a density of  $1.42\text{gcm}^{-3}$ . Calculate the molarity of concentrated nitric acid
2. Concentrated ammonia solution is 30%w/w and has a specific gravity of  $0.88\text{gcm}^{-3}$ . Calculate the molarity of concentrated ammonia

**(b) Diluting a more concentrated solution.**

Dilute standard solutions can be prepared from more concentrated standard solutions in a measured way;

In dilution, it is only the volume of the solvent that increases; the amount of the substance (number of moles) in the final less concentrated solution remains the

same. The number of moles of the solute before dilution equals to the number of moles of solute after dilution.

**For example:**

(a) *Preparing 250cm<sup>3</sup> of 0.1M hydrochloric acid from 2.0M hydrochloric acid*

The volume of concentrated acid to be diluted to 250cm<sup>3</sup> is first calculated from;

$$\frac{M_{conc} \times V_{conc}}{1000} = \frac{M_{dilute} \times V_{dilute}}{1000}$$

Or

$$M_C \times V_C = M_D \times V_D$$

$$\frac{2 \times V_{conc}}{1000} = \frac{0.1 \times 250}{1000}$$

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

Therefore 12.5cm<sup>3</sup> of 2M *HCl* are dissolved in a little water and the solution diluted up to the 1000cm<sup>3</sup> mark.

(b) Suppose you are provided with a stock solution of 11.6M *HCl*. How would you prepare a solution of 0.05M *HCl* solution?

11.6 moles of *HCl* are contained in 1000cm<sup>3</sup> of solution

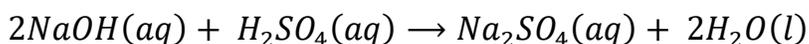
$$\begin{aligned} 0.05 \text{ moles of } HCl \text{ are contained in } & \left(\frac{1000}{11.6} \times 0.05\right) \text{ cm}^3 \text{ of solution} \\ & = 4.31\text{cm}^3 \end{aligned}$$

Therefore 4.31cm<sup>3</sup> of the stock solution are dissolved in a little water and the solution diluted up to the 1000cm<sup>3</sup> mark.

**Sample questions:**

1. A 25.0cm<sup>3</sup> solution of sulphuric acid was diluted to 500cm<sup>3</sup>. 25.0cm<sup>3</sup> of 1M Sodium hydroxide Solution required 20.0cm<sup>3</sup> of the diluted sulphuric acid for complete neutralization.

(a) Write the equation for the reaction.



(b) Calculate

(i) The concentration of the diluted acid

$1000\text{cm}^3$  contain 1 mole of  $\text{NaOH}$

$25\text{cm}^3$  contain  $\left[\frac{1}{1000} \times 25\right]$  moles of  $\text{NaOH}$

$=0.025$  moles of  $\text{NaOH}$

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

$\therefore$  Number of moles of acid reacted  $=\frac{1}{2} \times 0.025$

$=0.0125$  moles

$20\text{cm}^3$  contain 0.0125 moles of  $\text{H}_2\text{SO}_4$

$1000\text{cm}^3$  contain  $\frac{0.0125}{20} \times 1000$

$=0.625\text{M}$ .

**$\therefore$  concentration of diluted sulphuric acid = 0.625M**

(ii) The concentration of the original concentrated sulphuric acid solution.

$1000\text{cm}^3$  contain 0.625 moles of acid

$500\text{cm}^3$  contain  $\frac{0.625}{1000} \times 500$

$=0.3125$  moles of acid.

$25\text{cm}^3$  contain 0.3125 moles of acid.

$1000\text{cm}^3$  contain  $\frac{0.3125}{25} \times 1000$  moles

$=12.5\text{M}$

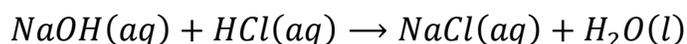
2.  $12\text{cm}^3$  of  $\text{XM}$  hydrochloric acid was diluted to  $250\text{cm}^3$ .  $26\text{cm}^3$  of the resultant solution reacted with  $20\text{cm}^3$  of  $0.1\text{M}$  sodium hydroxide solution. Determine the value of  $\text{X}$ .

Moles of NaOH that reacted

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

$$\begin{aligned} 20\text{cm}^3 \text{ of NaOH contain } & \left( \frac{0.1}{1000} \times 20 \right) \text{ mole} \\ & = 0.002 \text{ moles of NaOH} \end{aligned}$$

Moles of HCl that reacted



Mole ratio of NaOH:HCl = 1:1

Moles of HCl that reacted =  $0.002$

$26\text{cm}^3$  of diluted HCl contain  $0.002$  moles

$$250\text{cm}^3 \text{ of diluted HCl contain } \left( \frac{0.002}{26} \times 250 \right) \text{ moles}$$

Moles of acid in  $12\text{cm}^3$  of original acid = moles of acid in diluted  $250\text{cm}^3$  of solution

$$\text{Moles of acid in } 12\text{cm}^3 \text{ of original acid} = \left( \frac{0.002}{26} \times 250 \right) \text{ moles}$$

$$12\text{cm}^3 \text{ of original acid contain } \left( \frac{0.002}{26} \times 250 \right) \text{ moles}$$

$$\begin{aligned} 1000\text{cm}^3 \text{ of the original acid contain } & \left( \frac{0.002 \times 250 \times 1000}{26 \times 12} \right) \text{ moles} \\ & = 2 \text{ moles} \end{aligned}$$

$$\therefore \text{X} = 2$$

**Hint:**

Solution	Specific density	%w/w	Molar concentration
Hydrochloric acid	1.18	36	11.6
Nitric acid	1.42	70	15.8
Sulphuric acid	1.84	98	18.4
Conc. Ammonia	0.88	30	15.5

**EXERCISE:**

1. Calculate the concentration of a 36% hydrochloric acid solution of density  $1.18\text{gcm}^{-3}$
2. (a) Calculate the concentration of stock solution of nitric acid which has a mass density of  $1.50\text{gcm}^{-3}$  and has a purity of 70%  
  
(b). Calculate the volume of the stock solution that would be diluted to make 5 litres of 2M nitric acid.
3. Calculate ;  
(a) The volume of 2M sulphuric acid that would be diluted to make  $500\text{cm}^3$  of 0.8M sulphuric acid.  
(b) The volume of water that must be added to  $50\text{cm}^3$  of 2M sulphuric acid to make 0.15M sulphuric acid.
4.  $20\text{cm}^3$  of 2M sodium hydroxide solution is diluted to  $500\text{cm}^3$  with water.  
(a) What is the molarity of the dilute solution? (**Ans. Molarity=0.08M**)  
(b) Calculate the volume of water that must be added to  $100\text{cm}^3$  of 2M sodium hydroxide solution to make it 0.25M sodium hydroxide. (**Ans. Volume of water added=700cm<sup>3</sup>**)  
(c) Calculate the volume of 1.5M sulphuric acid that would be required to make 500ml of 0.2M sulphuric acid. (**volume of 1.5M= 66.67cm<sup>3</sup>**)

5. A  $50.0\text{cm}^3$  sample of sulphuric acid was diluted to  $1\text{ dm}^3$ . A sample of the diluted sulphuric acid was analyzed by titrating with aqueous sodium hydroxide. In the titration,  $25.0\text{cm}^3$  of  $1.0\text{mol dm}^{-3}$  aqueous sodium hydroxide required  $20.0\text{cm}^3$  of diluted sulphuric acid for complete neutralization.

(a) Write the equation for the reaction.

(b) Calculate;

(i) The concentration of the diluted acid. **(Ans.0.625M)**

(ii) The concentration of the original concentrated acid solution. **(Ans.12.5M)**

### ***APPLICATIONS OF STANDARD SOLUTION:***

It is applied in volumetric analysis in;

(a) **Standardization of a solution.**

Standardization refers to the process of determining the concentration of solution by titrating it with a solution of known concentration (primary standard.)

(b) **Standardization of a substance** by another substance which is not a primary standard, which has been standardized by a primary standard. E.g.

Standardizing sodium hydroxide by hydrochloric acid, which has been standardized by anhydrous sodium carbonate.

(c) Determining the basicity of an acid, percentage purity of a substance, relative atomic mass of a metal in a compound, reaction ratio e.t.c

### ***CLASSIFICATION/ TYPES OF REACTIONS IN VOLUMETRIC ANALYSIS***

There are basically three types of titration namely;

(a) **Acid-base titration.** (Neutralization reaction)

This is a titration in which a strong or weak acid is titrated against a strong or weak base till end point using either phenolphthalein or methyl orange indicator.

Acid-base titration can be classified into 2 or 3 types;

- (i) Single indicator acid-base titration
- (ii) Double indicator acid-base titration.
- (iii) Back titration.

(b) **Redox titrations** (oxidation-reduction reactions)

These include all reactions involving change of oxidation numbers or transfer of electron among the reacting substances.

(c) Precipitation reactions.

This depends upon the combination of ions to form a simple precipitate as in the titration of silver ions with the solution of a chloride.

***PRESENTATION AND TREATMENT OF RESULTS***

Consider the following example;

You are provided with the following;

**FA1**, which is approximately 0.1M hydrochloric acid

**Solid R**, which is sodium tetra borate decahydrate ( $Na_2B_4O_7 \cdot 10H_2O$ )

You are required to;

- (i) Prepare a standard solution of sodium tetra borate and use it to;
- (ii) Standardize hydrochloric acid in **FA1**

Procedure;

Weigh accurately 4.8g of **solid R** into a beaker and add about 100cm<sup>3</sup> of distilled water shake well or stir to dissolve. Transfer all the content into a 250cm<sup>3</sup> volumetric flask and make it to the mark with distilled water. Label the resultant solution **FA2**.

Pipette 20.00cm<sup>3</sup> (or 25.00cm<sup>3</sup>) of **FA2** into a clean conical flask. Add 2-3 drops of methyl orange indicator and titrate it with **FA1** from the burette until end point.

Repeat the titration until you obtain consistent results. Record your results in the table below.

Results:

Mass of beaker +solid R =114.8g

Mass of beaker =110.0g

**Record to at least 1 dp**

Mass of solid R =4.8g

Volume of pipette used=25.00cm<sup>3</sup>

Final burette reading (cm <sup>3</sup> )	24.00	47.50	33.60
Initial burette reading (cm <sup>3</sup> )	0.00	24.00	10.00
Volume of <b>FA1</b> used (cm <sup>3</sup> )	24.00	23.50	23.60

**Note the following about the table**

- (i) Each value in the table must be written to two decimal places, with the final digit being zero.
- (ii) Different initial readings may be used.
- (iii). The final burette reading must be greater than the initial burette reading.
- (iv) Each initial reading must be correctly subtracted from the final reading.
- (v). There should be consistency in at least two titre values.

Questions:

Volume of **FA1** used in calculating the average.

23.50 and 23.60cm<sup>3</sup>

*(Only two titre values from the table should be chosen for calculating the average volume and should not differ by  $\mp 0.10\text{cm}^3$ )*

Average of **FA1** used.

$$\frac{23.50 + 23.60}{2} = 23.55\text{cm}^3 \text{ (accuracy = true value } \mp 0.5)$$

(a) Calculate the number of moles of **FA2** that reacted with **FA1**

**Note: All calculations should be done from first principles.**

*Molar mass of  $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$*

$$\begin{aligned} &= (23 \times 2) + (10.8 \times 4) + (16 \times 7) + (10 \times 18) \\ &= 381.2\text{g} \end{aligned}$$

*381.2g of Borax contain 1 mole*

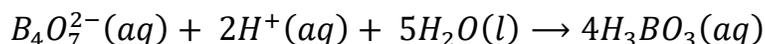
$$\begin{aligned} 4.8\text{g of Borax contain } &\left(\frac{1 \times 4.8}{381.2}\right) \text{ moles} \\ &= 0.0126 \text{ moles} \end{aligned}$$

*250cm<sup>3</sup> of **FA2** contain 0.0126 moles*

$$\begin{aligned} 25\text{cm}^3 \text{ of **FA2** contain } &\left(\frac{0.0126 \times 25}{250}\right) \text{ moles} \\ &= 0.00126 \text{ moles of **FA2**} \end{aligned}$$

(b) Determine the number of moles of **FA1** that reacted with **FA2**

*Write the equation for the reaction if not given, then use the reacting ratio*



*1 mole of **FA2** reacts with 2 moles of **FA1***

$$\begin{aligned} 0.00126 \text{ moles **FA2** react with } &\left(\frac{2 \times 0.00126}{1}\right) \text{ moles of **FA1**} \\ &= 0.00252 \text{ moles of **FA1**} \end{aligned}$$

**$\therefore$  Moles of **FA1** that reacted with **FA2** = 0.00252**

(c) Calculate the concentration of hydrochloric acid in **FA1** in moles per litre.

*23.55cm<sup>3</sup> of **FA1** contain 0.00252 moles of **HCl***

$$1000\text{cm}^3 \text{ of FA1 contain } \left( \frac{0.00252 \times 1000}{23.55} \right) \text{ moles of HCl}$$

$$= 0.106 \text{ moles of HCl}$$

**$\therefore$  Concentration of HCl in FA1 = 0.106M**

*Note: The final answer will depend on the accuracy of the titre value.*

## QUESTIONS

- 25cm<sup>3</sup> of 0.1M solution of sodium hydroxide needed 20cm<sup>3</sup> of hydrochloric acid to reach end point in the presence of phenolphthalein indicator. Calculate the concentration of hydrochloric acid in grams per litre. (*H* = 1; *Cl* = 35.5)
- 1.6g of an acid  $C_2O_4H_2 \cdot nH_2O$  was dissolved in water and made up to the 250cm<sup>3</sup> mark in a volumetric flask. 25.0cm<sup>3</sup> portions of this solution was titrated with 0.1M sodium hydroxide solution using phenolphthalein indicator. 25.10cm<sup>3</sup> of the base was needed for complete neutralization. Calculate;
  - The RMM of the acid. (**Ans RMM  $\approx$  127.5**)
  - The value of *n* in the formula  $C_2O_4H_2 \cdot nH_2O$ . (*C* = 12; *H* = 1; *O* = 16)(***n* = 2**)
- 1.55g of an acid  $C_nH_{2n}(COOH)_2$  was dissolved in water and the solution made up to 250cm<sup>3</sup>. 25.0cm<sup>3</sup> of this solution required 23.50cm<sup>3</sup> of a 0.1M sodium hydroxide solution for complete neutralization.
  - Write the equation for the reaction that took place.
  - Calculate;
    - The molar concentration of the acid. (**Ans; MM  $\approx$  132**)
    - The molecular mass of the acid. Hence determine the value of *n* and write the structural formula of the acid.(**Ans *n* = 3**)
- 20.0cm<sup>3</sup> of a 0.02M solution of sodium hydroxide was added to 30.0cm<sup>3</sup> of 0.025M sulphuric acid. Calculate;
  - The molar concentration of the hydrogen ions in the initial sulphuric acid.
  - Concentration of the hydrogen ions in the resultant solution.
  - PH of the resultant solution.

**(UNEB 1998 P1 Qtn. 5)**

5. 24.0cm<sup>3</sup> of a solution containing 14.0g per litre of potassium hydroxide required 25.00cm<sup>3</sup> of 0.12M phosphoric acid for complete neutralization  
(K = 39, O = 16, H = 1)
- (a) Calculate the molar ratio in which phosphoric acid reacts with potassium hydroxide.
- (b) Write the equation for the reaction.
6. A 25.0cm<sup>3</sup> solution of sulphuric acid was diluted to 500cm<sup>3</sup>. 25.0cm<sup>3</sup> of 1M Sodium hydroxide Solution required 20.0cm<sup>3</sup> of the diluted sulphuric acid for complete neutralization.
- (a) Write the equation for the reaction.
- (b) Calculate;
- (i) The concentration of the diluted acid
- (ii) The concentration of the original concentrated sulphuric acid solution.
7. A 25.0cm<sup>3</sup> solution of sulphuric acid was diluted to 500cm<sup>3</sup>. 25.0cm<sup>3</sup> of 1M Sodium hydroxide Solution required 20.0cm<sup>3</sup> of the diluted sulphuric acid for complete neutralization.
- (a) Write the equation for the reaction.
- $$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
- (b) Calculate
- (i) The concentration of the diluted acid
- 1000cm<sup>3</sup> of solution contain 1 mole of NaOH
- 25cm<sup>3</sup> of solution contain  $(\frac{1}{1000} \times 25)$  moles of NaOH
- =0.025 moles
- 2 moles of NaOH react with 1 mole of H<sub>2</sub>SO<sub>4</sub>

Number of moles of  $H_2SO_4$  that reacted =  $\frac{1}{2} \times 0.025 = 0.0125$  moles

20  $cm^3$  of solution contain 0.0125 moles of  $H_2SO_4$

1000  $cm^3$  of solution contain  $(\frac{0.0125}{20} \times 1000)$  moles of  $H_2SO_4$

$$= 0.625M$$

(ii) The concentration of the original concentrated sulphuric acid solution.

1000  $cm^3$  of solution contain 0.625 moles of acid.

500  $cm^3$  of solution contain  $(\frac{0.625}{1000} \times 500)$  moles of the acid

$$= 0.3125 \text{ Moles of acid.}$$

25  $cm^3$  of original acid contained 0.3125 moles

1000  $cm^3$  of solution contain  $(\frac{0.3125}{25} \times 1000)$  moles

$$= 12.5M$$

**$\therefore$  Concentration of the original concentrated sulphuric acid solution = 12.5M**

8. 2.5g of hydrated sodium carbonate,  $Na_2CO_3 \cdot XH_2O$ , were dissolved in 250  $cm^3$  of water. 25  $cm^3$  of the resultant solution required 17.5  $cm^3$  of 0.1M hydrochloric acid to neutralize it completely. Calculate;
  - (a) The molarity of sodium carbonate solution.
  - (b) The relative molecular mass  $Na_2CO_3 \cdot XH_2O$ .
  - (c) The value of X in  $Na_2CO_3 \cdot XH_2O$  ( $Na = 23, C = 12, O = 16, H = 1$ )
9. 5.34g of a salt of formula  $M_2SO_4$  (where M is a metal) were dissolved in water. The sulphate ion was precipitated by adding excess barium chloride solution and 4.66g of barium sulphate were obtained. ( $Ba = 137, S = 32, O = 16$ )

Calculate;

- (a) The number of moles of  $M_2SO_4$  present in the solution.
- (b) The formula mass of  $M_2SO_4$
- (c) The relative atomic mass of M.

### *TYPES OF TITRATIONS*

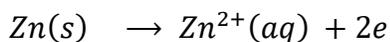
-  Redox titrations.
-  Back titrations.
-  Double indicator titrations.

### *REDOX REACTIONS*

A redox reaction refers to a reaction in which both reduction and oxidation reactions occur simultaneously.

#### *OXIDATION*

This is defined as the loss of electron(s) from a chemical substance (or species.)

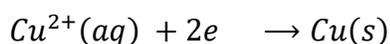


(0)                    (+2)

Here, there is an increase in oxidation number (state) as indicated in the brackets below the species.

#### *REDUCTION*

This is defined as the gain of electron(s) by a chemical substance (or species.)



(+2)                    (0)

Here, there is a decrease in oxidation number (state) as indicated in the brackets below the species.

#### *Note;*

The electrons lost by zinc must be gained by copper(II) ions; therefore zinc must be a reducing agent while copper(II) ions must be oxidizing agent.

(i) **Common Oxidizing agents**

- ❖ Potassium manganate(VII),  $KMnO_4$
- ❖ Potassium dichromate(VI),  $K_2Cr_2O_7$
- ❖ Manganese(IV) oxide,  $MnO_2$
- ❖ Hydrogen peroxide,  $H_2O_2$
- ❖ Nitric acid,  $HNO_3$
- ❖ Chlorates,  $ClO_3$
- ❖ Bromates,  $BrO_3$
- ❖ Iodates,  $IO_3$  e.t.c.

(ii) **Common Reducing agents**

- ❖ Ethanedioates(oxalates),  $C_2O_4^{2-}$
- ❖ Iodides,  $I^-$
- ❖ Nitrites,  $NO_2^-$
- ❖ Tin(II) ions,  $Sn^{2+}$
- ❖ Iron(II) salts,  $Fe^{2+}$
- ❖ Thiosulphates,  $S_2O_3^{2-}$
- ❖ Sulphites,  $SO_3^{2-}$  e.t.c

Study the tables below

Oxidizing agent	Formula	Oxidizing species	Reduced to
Potassium manganate(VII)	$KMnO_4$	$MnO_4^-$	$Mn^{2+}$
Potassium dichromate(VI)	$K_2Cr_2O_7$	$Cr_2O_7^{2-}$	$Cr^{3+}$
Potassium Iodate	$KIO_3$	$IO_3^-$	$I_2$
Manganese(IV) oxide	$MnO_2$	$MnO_2$	$Mn^{2+}$
Lead(IV) oxide	$PbO_2$	$PbO_2$	$Pb^{2+}$
Hydrogen peroxide	$H_2O_2$	$H_2O_2$	$H_2O$

Potassium persulphate	$K_2S_2O_8$	$S_2O_8^{2-}$	$SO_4^{2-}$
-----------------------	-------------	---------------	-------------

The table below shows common reducing agents and the species to which they are oxidized.

Reducing agent	Formula	Active species	Oxidized to
Potassium iodide	$KI$	$I^-$	$I_2$
Oxalates e.g. Sodium oxalate Oxalic acid	$Na_2C_2O_4$ $H_2C_2O_4$	$C_2O_4^{2-}$	$CO_2$
Hydrogen peroxide	$H_2O_2$	$H_2O_2$	$O_2$ and $H^+$
Sulphites e.g. Sodium sulphite	$Na_2SO_3$	$SO_3^{2-}$	$SO_4^{2-}$
Thiosulphates e.g. Sodium thiosulphate	$Na_2S_2O_3$	$S_2O_3^{2-}$	$S_4O_6^{2-}$

### **BALANCING REDOX REACTIONS**

- ✓ A redox reaction is made up of two parts called half equations or reactions. i.e. oxidation half equation and reduction half equation.
- ✓ These half equations are written separately, balanced and then added together to form the overall Stoichiometric equation of reaction (overall redox reaction equation.)
- ✓ To balance redox equations, one should identify the medium under which the reaction takes place .i.e. acidic, alkaline or neutral.

#### **Redox reactions in acidic medium.**

The following rules apply to acidic solutions.

- Identify the product for each reacting species. E.g.  
 $MnO_4^-$  is reduced to  $Mn^{2+}$   
 $Cr_2O_7^{2-}$  is reduced to  $Cr^{3+}$
- Write down the separate oxidation and reduction half equations.
- Balance the number of atoms of each element in each half equation.

- Balance the oxygen atoms by adding appropriate number of water molecule to the oxygen deficient side.
- Balance the hydrogen atoms by adding hydrogen ions on the water deficient side.
- Balance each half equation in terms of charge by adding appropriate number of electrons on the electron deficient side.
- Finally obtain the overall Stoichiometric equation of reaction by adding the two half equations but only after balancing the number of electrons in the two half equations by multiplying by suitable integers.

### **EXAMPLES**

#### **1) Reaction between $MnO_4^-$ and $Fe^{2+}$ in acidic medium**

Write half equations and overall reaction equation for the reaction between  $Mn^{2+}$  and  $MnO_4^-$  in acidic medium.

Reduction half equation.



Add 4 molecules of water of the R.H.S to balance oxygen atoms.



Add 8 hydrogen ions on the L.H.S to balance hydrogen atoms.



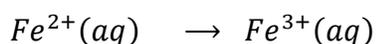
Total charge on L.H.S =  $-1 + 8 = +7$

Total charge on R.H.S =  $+2$

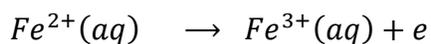
Add 5 electrons on the L.H.S so that each side remains with a charge of  $+2$



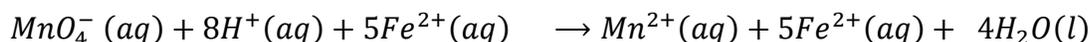
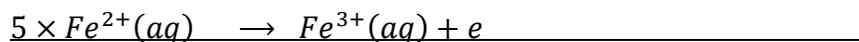
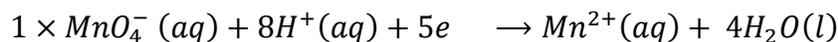
Oxidation half equation.



Since atoms are balanced, add one electron on the R.H.S so that each side remains with a charge of  $+2$



### Overall reaction equation



### **Exercise**

(a) Write half equations and overall reaction equation for each of the following reactions ;

- i) Reaction between  $\text{Cr}_2\text{O}_7^{2-}$  and  $\text{Fe}^{2+}$  in acidic medium.
- ii) Reaction between  $\text{MnO}_4^- (\text{aq})$  and sodium nitrite in acidic medium.
- iii) Reaction between dichromate ions and sulphite ions in acidic medium.
- iv) Reaction between  $\text{MnO}_4^- (\text{aq})$  and hydrogen peroxide in acidic medium.
- v) Reaction between  $\text{MnO}_4^- (\text{aq})$  and oxalate ions in acidic medium.

### **Titration with Potassium manganate(VII) and potassium dichromate**

- Manganate(VII) ions is about the most powerful oxidant, therefore very useful in volumetric analysis particularly since it requires no indicator.
- In acidic medium, it is reduced from  $\text{Mn}^{7+}$  to  $\text{Mn}^{2+}$  by reducing agents, its colour changes from purple to colourless or faint pink due to reduction from  $\text{Mn}^{7+}$  to  $\text{Mn}^{2+}$ .



**(Purple)** **( Colourless)**

- In alkaline medium,  $\text{MnO}_4^- (\text{aq})$  is reduced to  $\text{MnO}_2$  .i.e. the purple solution turns to brown due to reduction from  $\text{Mn}^{7+}$  to  $\text{Mn}^{4+}$



**(Purple)** **(brown)**

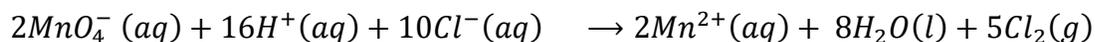
$\text{MnO}_4^- (\text{aq})$  Therefore has an advantage of being a self indicator.

#### **N. B**

- The manganate(VII) solution should be added slowly to allow reaction to occur. It is advisable to read the upper meniscus, as it is difficult to read the lower meniscus.
- Only sulphuric acid should be used to acidify the solution. Nitric acid is unsuitable because itself, nitric acid also can be an oxidizing agent and therefore, interferes with the products of the reaction.
- Hydrochloric acid is also unsuitable because it is oxidized by  $\text{MnO}_4^- (\text{aq})$  to chlorine.

#### **Questions.**

1. (a) Iron(II) sulphate is normally used to standardize a solution of potassium manganate(VII) acidified using 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.
    - Hydrochloric acid is oxidized by potassium manganate(VII) to chlorine.



(b) 25cm<sup>3</sup> of an acidified solution of 0.02M potassium manganate(VII) reacted completely with 25cm<sup>3</sup> of sodium nitrite.

- (i) Write the equation for the reaction that took place.
  - (ii) Calculate the concentration of sodium nitrite in grams per litre.  
(Na = 23; H = 1; N = 14; O = 16)
2. 1.35g of solid oxalate (COOH)<sub>2</sub>M<sub>x</sub> were weighed into a 250cm<sup>3</sup> volumetric flask, dissolved in water and the resultant solution made to the mark. 25.0cm<sup>3</sup> of this solution was acidified with 2M sulphuric acid and the mixture warmed to 60°C and titrated with 0.02M potassium manganate(VII) solution from the burette. The volume of manganate(VII) required for complete reaction was 20.20cm<sup>3</sup>.

Calculate;

- (a) The number of moles of MnO<sub>4</sub><sup>-</sup> (aq) that reacted.
  - (b) The number of moles of oxalate in 250cm<sup>3</sup>.
  - (c) the relative molecular mass of the oxalate salt.
  - (d) The value of x in (COOH)<sub>2</sub>M<sub>x</sub> if M = 23
3. 0.9875g of an impure potassium manganate(VII) was dissolved in water to make 250cm<sup>3</sup> of solution. When 20.0cm<sup>3</sup> 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 250cm<sup>3</sup> of solution. 24.40cm<sup>3</sup> of the sodium ethanedioate solution was used. (Na<sub>2</sub>C<sub>2</sub>O<sub>4</sub> = 134 and KMnO<sub>4</sub> = 158)

- (a) Write ionic equation for the reaction between sodium ethanedioate and potassium manganate(VII).
- (b) Determine the molar concentration of manganate(VII) ions. (**Ans = 0.0244M**)
- (c) Calculate the percentage purity of potassium manganate(VII). (**Ans = 97.6%**)
- (d) Name one compound which is a common impurity in potassium manganate(VII)

- **Manganese(IV) oxide.**

4. (a) Potassium manganate(VII) is not used as a primary standard in volumetric analysis and has to be standardized.
- What is meant by the term primary standard?
  - State three characteristics of a primary standard.
  - Explain why potassium manganate(VII) is not used as a primary standard.
  - Name one substance, other than Ethanedioic acid, which can be used as a primary standard for potassium manganate(VII)

(b) Explain why hydrochloric acid is not usually used to acidify potassium manganate(VII).

(c) Acidified potassium manganate(VII) reacts with Ethanedioic acid.

Write;

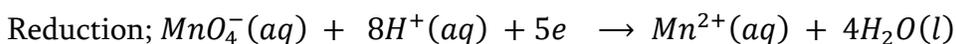
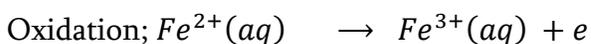
- The half equations for the reaction.
  - The overall equation for the reaction.
- (d) 20.00cm<sup>3</sup> of a 0.01M manganate(VII) ion solution required exactly 16.65cm<sup>3</sup> of a solution containing 4.8g/l of an oxalate(COO<sup>-</sup>X<sup>+</sup>)<sub>2</sub>.2H<sub>2</sub>O. Calculate the atomic mass of X.
5. 5g of solid sodium permanganate were dissolved in 250cm<sup>3</sup> of distilled water to make a solution. 20cm<sup>3</sup> of this solution was acidified by an equal volume of 1M sulphuric acid which when titrated with a 0.02M potassium oxalate solution, 31.60cm<sup>3</sup> of the solution was required.

- (a) Write the half equations for the reactions that occur and hence deduce the overall redox reaction equation.
- (b) Calculate the percentage of manganese in the solid.
6. A solution of a metal dichromate contains 3g of  $MCr_2O_7$  per litre of solution. 25cm<sup>3</sup> of this solution when acidified required 42.7cm<sup>3</sup> of a solution containing 2g of  $Fe^{2+}$  ions per litre for complete reaction.
- (a) Write the half equations for the reaction that occur and hence deduce the overall redox reaction equation.
- (b) Calculate the relative atomic mass of M ( $Fe = 56; Cr = 52; O = 16$ )
7. 25.0cm<sup>3</sup> of acidified hydrogen peroxide solution required 19.00cm<sup>3</sup> of 0.02M potassium manganate(VII) solution for complete reaction.
- (a) Write down the separate half equations and hence the overall redox equation of the reaction.
- (b) Calculate the concentration of hydrogen peroxide solution in moles per litre.
- (c) Briefly explain why potassium manganate(VII) solution is not a good primary standard.
- (d) Explain why potassium manganate(VII) solution is commonly used in titration despite being a poor primary standard.

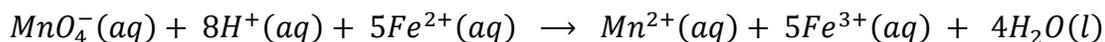
## 2) Analysis of a mixture of $Fe^{2+}$ and $Fe^{3+}$ ions

1. 25cm<sup>3</sup> of a solution containing a mixture of iron(II) and iron(III) ions was acidified with dilute sulphuric acid and titrated with 0.02M potassium manganate(VII). 20cm<sup>3</sup> of potassium manganate(VII) was needed to reach end point. A second fresh 25cm<sup>3</sup> of the mixture was first reduced with zinc powder and then on titration with 0.02M potassium manganate(VII) solution, 24.00cm<sup>3</sup> was required to reach end point.

### Half equations;



Overall equation;



Calculate;

(a) The concentration of iron(II) in the mixture in moles per dm<sup>3</sup>

1000cm<sup>3</sup> of  $\text{MnO}_4^-(aq)$  contain 0.02 moles

20cm<sup>3</sup> of  $\text{MnO}_4^-(aq)$  contain  $\left(\frac{0.02}{1000} \times 20\right)$  moles

= 0.0004 moles of  $\text{MnO}_4^-(aq)$

1 mole of  $\text{MnO}_4^-(aq)$  reacts with 5 moles of  $\text{Fe}^{2+}(aq)$

0.0004 moles of  $\text{MnO}_4^-(aq)$  react with  $\left(\frac{5}{1} \times 0.0004\right)$  moles of  $\text{Fe}^{2+}$

=0.002 moles of  $\text{Fe}^{2+}$

25cm<sup>3</sup> of the mixture contain 0.002 moles of  $\text{Fe}^{2+}$

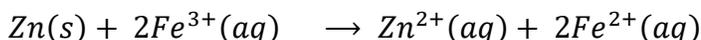
1000cm<sup>3</sup> of the mixture contain  $\left(\frac{0.002}{25} \times 1000\right)$  moles of  $\text{Fe}^{2+}$

= 0.08M

∴ Concentration of  $\text{Fe}^{2+}$  in the mixture = 0.08M

(b) The concentration of iron(III) in the mixture in moles per dm<sup>3</sup>

Zinc powder reduces  $\text{Fe}^{3+}$  in the mixture to  $\text{Fe}^{2+}$



Therefore, in the second titration, the  $\text{MnO}_4^-(aq)$  react with  $\text{Fe}^{2+}$  from the reduction of  $\text{Fe}^{3+}$  originally present in the mixture and the equation is similar to that in (a) above.

1000cm<sup>3</sup> of  $\text{MnO}_4^-(aq)$  contain 0.02 moles

24cm<sup>3</sup> of  $\text{MnO}_4^-(aq)$  contain  $\left(\frac{0.02}{1000} \times 24\right)$  moles

= 0.00048 moles of  $\text{MnO}_4^-(aq)$

1 mole of  $\text{MnO}_4^-(aq)$  reacts with 5 moles of  $\text{Fe}^{2+}(aq)$

0.00048 moles of  $\text{MnO}_4^-(aq)$  react with  $\left(\frac{5}{1} \times 0.00048\right)$  moles of  $\text{Fe}^{2+}$

=0.0024 moles of  $\text{Fe}^{2+}$

25cm<sup>3</sup> of the reduced mixture contain 0.0024 moles of  $Fe^{2+}$

$$1000\text{cm}^3 \text{ of the reduced mixture contain } \left(\frac{0.0024}{25} \times 1000\right) \text{ moles of } Fe^{2+} \\ = 0.096M$$

Total concentration of  $Fe^{2+} = 0.096M$

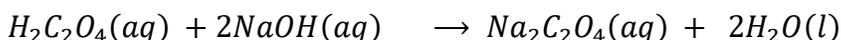
$$\therefore \text{Concentration of } Fe^{3+} \text{ in the mixture} = (0.096 - 0.08) \\ = 0.16M$$

***Analysis of a mixture of ethanedioic acid (oxalic acid) and sodium ethanedioate (sodium oxalate.)***

2. 20cm<sup>3</sup> of a mixture of oxalic acid and sodium oxalate solution was titrated with 0.1M sodium hydroxide using phenolphthalein indicator. 10.0cm<sup>3</sup> of the base was required to reach end point. Another fresh 20cm<sup>3</sup> of the mixture was mixed with 20cm<sup>3</sup> of 2M sulphuric acid in a conical flask and the mixture heated to about 50-60°C and immediately titrated with 0.02M potassium manganate(VII) solution. 25.50cm<sup>3</sup> of potassium manganate(VII) solution was required to reach the end point.

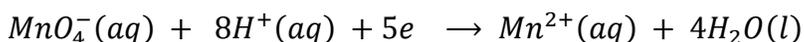
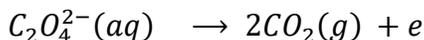
**Theory;**

In the first titration, only oxalic acid reacts with sodium hydroxide as

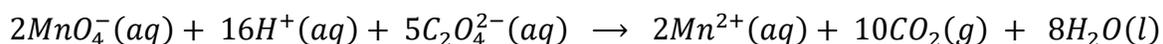


In the second titration, the  $C_2O_4^{2-}$  ions from both the oxalic acid and sodium oxalate are all oxidized by acidified  $MnO_4^-$  ions as

**Half equations;**



**Overall equation;**



Calculate;

- (a) The concentration of oxalic acid  $H_2C_2O_4 \cdot 2H_2O$  in g/l in the mixture.

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

10.0cm<sup>3</sup> of NaOH contain  $\left(\frac{0.1}{1000} \times 10.0\right)$  moles

$$=0.001 \text{ moles of NaOH}$$

2 moles of NaOH react with 1 mole of oxalic acid

0.001mole of NaOH react with  $\left(\frac{1}{2} \times 0.001\right)$  moles of oxalic acid

$$= 0.0005 \text{ moles of oxalic acid.}$$

20cm<sup>3</sup> of the mixture contain 0.0005 moles of oxalic acid.

1000cm<sup>3</sup> of the mixture contain  $\left(\frac{0.0005}{20} \times 1000\right)$  moles of oxalic acid

$$= 0.025\text{M}$$

RFM of  $H_2C_2O_4 \cdot 2H_2O = (1 \times 2) + (12 \times 2) + (16 \times 4) + (2 \times 18)$

$$= 126$$

1mole of oxalic acid weighs 126g

0.025mles of oxalic acid weigh  $\left(\frac{126}{1} \times 0.025\right)$

$$= 3.15\text{g/l}$$

$\therefore$  Concentration of oxalic acid = 3.15g/l

(b) The concentration of sodium oxalate in g/l in the mixture.

1000cm<sup>3</sup> of  $MnO_4^-$  contain 0.02 moles

25.50cm<sup>3</sup> of  $MnO_4^-$  contain  $\left(\frac{0.02}{1000} \times 25.50\right)$  moles

$$=0.00051 \text{ moles of } MnO_4^-$$

2moles of  $MnO_4^-$  react with 5 moles of  $C_2O_4^{2-}$

0.00051 moles of  $MnO_4^-$  react with  $\left(\frac{5}{2} \times 0.00051\right)$  moles of  $C_2O_4^{2-}$

$$= 0.001275 \text{ moles of } C_2O_4^{2-}$$

$\therefore$  20cm<sup>3</sup> of the mixture contains 0.001275 moles of total  $C_2O_4^{2-}$

1000cm<sup>3</sup> of the mixture contain  $\left(\frac{0.001275}{20} \times 1000\right)$  moles of  $C_2O_4^{2-}$

$$= 0.06375\text{mol/l}$$

Total concentration of  $C_2O_4^{2-} = 0.06375 \text{ mol/l}$

Therefore, concentration of  $C_2O_4^{2-}$  from sodium oxalate =  $0.06375 - 0.025$   
 $= 0.03875 \text{ mol/l}$

RFM of  $Na_2C_2O_4 = (23 \times 2) + (12 \times 2) + (16 \times 4) = 134$

1 mole of  $Na_2C_2O_4$  weighs 134g

0.03875 mole of  $Na_2C_2O_4$  weigh  $(\frac{134}{1} \times 0.03875)$   
 $= 5.1925 \text{ g/l}$

Therefore, concentration of sodium oxalate =  $5.1925 \text{ g/l}$

(c) The percentage of oxalic acid in the mixture.

$$\text{Percentage of oxalic acid} = \left( \frac{3.15}{(3.15+5.1925)} \times 100 \right)$$
$$= 37.76\%$$

(d) The percentage of sodium oxalate in the mixture

( $Na = 23, C = 12, O = 16, H = 1$ )

$$\text{Percentage of sodium oxalate} = 100 - 37.76$$
$$= 62.24\%$$

### **DOUBLE INDICATOR TITRATION.**

This is a titration technique used when determining the percentage composition by mass of a specific component in a given mixture.

#### **Types of mixtures**

- Complex mixtures. Here both components react with the acid in the same reaction. E.g.
  - ✓ A mixture of sodium hydroxide and sodium carbonate.
  - ✓ A mixture of sodium hydrogen carbonate and sodium carbonate.
  - ✓ A mixture of sodium hydroxide and sodium hydrogen carbonate.

Analysis of a mixture of sodium hydroxide and sodium carbonate by the continuous method.

25.0cm<sup>3</sup> of a solution containing a mixture of sodium hydroxide and sodium carbonate required 23.50cm<sup>3</sup> of 0.1M hydrochloric acid using phenolphthalein indicator for reaction to reach end point. The same solution was titrated further using methyl orange indicator and it required 10.7cm<sup>3</sup> of 0.1M hydrochloric acid for the reaction to reach end point. Calculate;

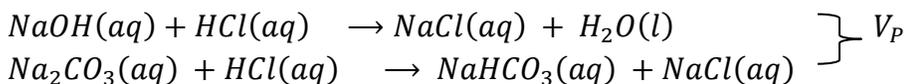
- (a) The concentration of sodium hydroxide in grams per litre.  
 (b) The concentration of sodium carbonate in grams per litre.

**Solution**

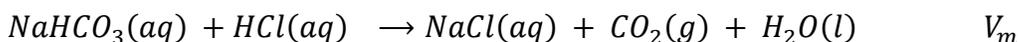
Theory:

The reaction of sodium hydroxide occurs in a single step while that of sodium carbonate occurs in two steps.

With phenolphthalein indicator, the sodium hydroxide is completely neutralized while sodium carbonate is half neutralized.



With methyl orange indicator, the sodium hydrogen carbonate formed is converted to sodium chloride and carbon dioxide.



Since only half of sodium carbonate is neutralized using phenolphthalein as indicator, then the other half is neutralized using methyl orange as indicator. Therefore the volume of the acid needed for complete neutralization of sodium carbonate is twice  $V_m$

$$\begin{aligned} \text{i.e. Volume of acid for complete neutralization of sodium carbonate} &= 2V_m \\ &= 2 \times 10.7 \\ &= 21.40\text{cm}^3 \end{aligned}$$

$$\begin{aligned} \text{Volume of acid needed for complete neutralization of NaOH} &= V_P - V_m \\ &= 23.50 - 10.70 \\ &= 12.80\text{cm}^3 \end{aligned}$$

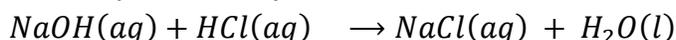
- (a) Moles of hydrochloric acid that reacted.

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

12.80cm<sup>3</sup> of HCl contain  $\left(\frac{0.1}{1000} \times 12.80\right)$  moles

= 0.00128 moles of Hydrochloric acid.

Moles of sodium hydroxide that reacted



1 mole of HCl reacts with 1 mole of NaOH

$$0.00128 \text{ moles of HCl react with } \left(\frac{1}{1} \times 0.00128\right) \text{ moles of NaOH}$$

$$= 0.00128 \text{ moles of NaOH}$$

Molarity of sodium hydroxide

25cm<sup>3</sup> of the mixture contain 0.00128 moles of NaOH

$$1000\text{cm}^3 \text{ of the mixture contain } \left(\frac{0.00128}{25} \times 1000\right)$$

$$= 0.0512\text{M}$$

Concentration NaOH in gl<sup>-1</sup>

$$\text{RFM of NaOH} = (23 \times 1) + (16 \times 1) + (1 \times 1)$$

$$= 40$$

1 mole of NaOH weighs 40g

$$0.0512 \text{ mole of NaOH weigh } \left(\frac{40}{1} \times 0.0512\right) \text{g}$$

$$= 2.048\text{gl}^{-1}$$

$$\therefore \text{Concentration of NaOH} = 2.048\text{gl}^{-1}$$

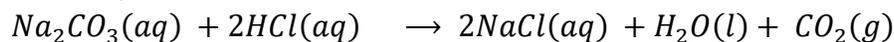
(b) Moles of hydrochloric acid that reacted.

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

$$21.40\text{cm}^3 \text{ of HCl contain } \left(\frac{0.1}{1000} \times 21.40\right) \text{ moles}$$

$$= 0.00214 \text{ moles of Hydrochloric acid.}$$

Moles of sodium carbonate that reacted



2 moles of HCl react with 1 mole of Na<sub>2</sub>CO<sub>3</sub>

$$0.00214 \text{ moles of HCl react with } \left(\frac{1}{2} \times 0.00214\right) \text{ moles of Na}_2\text{CO}_3$$

$$= 0.00107 \text{ moles of Na}_2\text{CO}_3$$

Molarity of sodium carbonate

25cm<sup>3</sup> of the mixture contain 0.00107 moles of Na<sub>2</sub>CO<sub>3</sub>

$$1000\text{cm}^3 \text{ of the mixture contain } \left(\frac{0.00107}{25} \times 1000\right)$$

$$= 0.0428\text{M}$$

Concentration Na<sub>2</sub>CO<sub>3</sub> in gl<sup>-1</sup>

$$\text{RFM of Na}_2\text{CO}_3 = (23 \times 2) + (12 \times 1) + (16 \times 3)$$

$$= 106$$

1 mole of Na<sub>2</sub>CO<sub>3</sub> weighs 106g

$$0.0428 \text{ moles of Na}_2\text{CO}_3 \text{ weigh } \left(\frac{106}{1} \times 0.0428\right) \text{g}$$

$$= 4.5368\text{gl}^{-1}$$

$$\therefore \text{Concentration of Na}_2\text{CO}_3 = 4.5368\text{gl}^{-1}$$

Analysis of a mixture of sodium hydroxide and sodium carbonate by the separate method.

25.0cm<sup>3</sup> of a mixture of sodium hydroxide and sodium carbonate required 18.70cm<sup>3</sup> of 0.05M sulphuric acid using phenolphthalein indicator.

Another 25.0cm<sup>3</sup> of the mixture required 23.50cm<sup>3</sup> of 0.05M sulphuric acid using methyl orange indicator.

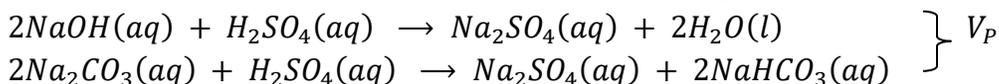
Determine the concentration of;

- (a) Sodium hydroxide in grams per litre.
- (b) Sodium carbonate in grams per litre.

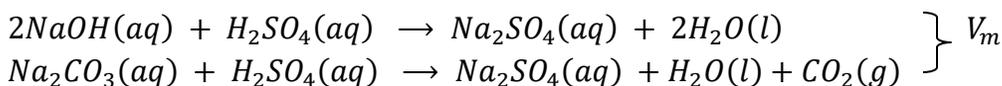
**Solution:**

Theory:

- With phenolphthalein indicator, sodium hydroxide is completely neutralized while sodium carbonate is half neutralized to sodium hydrogen carbonate.



- With methyl orange indicator, both sodium hydroxide and sodium carbonate are fully neutralized.



- Therefore, volume of the acid needed for complete neutralization of;

$$\begin{aligned} \color{red}{\oplus} \text{Na}_2\text{CO}_3 &= 2(V_m - V_p) \\ &= 2(23.50 - 18.70) \\ &= 9.60\text{cm}^3 \end{aligned}$$

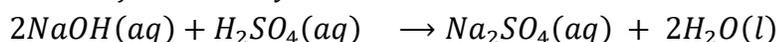
$$\begin{aligned} \color{red}{\oplus} \text{NaOH} &= 2V_p - V_m \\ &= (2 \times 18.70) - 23.50 \\ &= 13.90\text{cm}^3 \end{aligned}$$

- (a) Moles of sulphuric acid that reacted.

1000cm<sup>3</sup> of H<sub>2</sub>SO<sub>4</sub> contain 0.1 mole

$$\begin{aligned} 13.90\text{cm}^3 \text{ of H}_2\text{SO}_4 \text{ contain } &\left(\frac{0.05}{1000} \times 13.90\right) \text{ moles} \\ &= 0.000695 \text{ moles of H}_2\text{SO}_4. \end{aligned}$$

Moles of sodium hydroxide that reacted



1 mole of H<sub>2</sub>SO<sub>4</sub> reacts with 2 mole of NaOH

$$0.000695 \text{ moles of } H_2SO_4 \text{ react with } \left(\frac{2}{1} \times 0.000695\right) \text{ moles of NaOH}$$

$$= 0.00139 \text{ moles of NaOH}$$

Molarity of sodium hydroxide

25cm<sup>3</sup> of the mixture contain 0.00139 moles of NaOH

$$1000\text{cm}^3 \text{ of the mixture contain } \left(\frac{0.00139}{25} \times 1000\right)$$

$$= 0.0556M$$

Concentration NaOH in gl<sup>-1</sup>

$$RFM \text{ of NaOH} = (23 \times 1) + (16 \times 1) + (1 \times 1)$$

$$= 40$$

1 mole of NaOH weighs 40g

$$0.0556 \text{ mole of NaOH weigh } \left(\frac{40}{1} \times 0.0556\right) g$$

$$= 2.224\text{gl}^{-1}$$

$$\therefore \text{Concentration of NaOH} = 2.224\text{gl}^{-1}$$

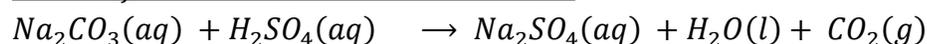
(b) Moles of sulphuric acid that reacted.

1000cm<sup>3</sup> of H<sub>2</sub>SO<sub>4</sub> contain 0.05 mole

$$9.60\text{cm}^3 \text{ of } H_2SO_4 \text{ contain } \left(\frac{0.05}{1000} \times 9.60\right) \text{ moles}$$

$$= 0.00048 \text{ moles of sulphuric acid.}$$

Moles of sodium carbonate that reacted



1 moles of H<sub>2</sub>SO<sub>4</sub> react with 1 mole of Na<sub>2</sub>CO<sub>3</sub>

$$0.00048 \text{ moles of } H_2SO_4 \text{ react with } \left(\frac{1}{1} \times 0.00048\right) \text{ moles of } Na_2CO_3$$

$$= 0.00048 \text{ moles of } Na_2CO_3$$

Molarity of sodium carbonate

25cm<sup>3</sup> of the mixture contain 0.00048 moles of Na<sub>2</sub>CO<sub>3</sub>

$$1000\text{cm}^3 \text{ of the mixture contain } \left(\frac{0.00048}{25} \times 1000\right)$$

$$= 0.0192M$$

Concentration Na<sub>2</sub>CO<sub>3</sub> in gl<sup>-1</sup>

$$RFM \text{ of } Na_2CO_3 = (23 \times 2) + (12 \times 1) + (16 \times 3)$$

$$= 106$$

1 mole of Na<sub>2</sub>CO<sub>3</sub> weighs 106g

$$0.0192 \text{ moles of } Na_2CO_3 \text{ weigh } \left(\frac{106}{1} \times 0.0192\right) g$$

$$= 2.0352\text{gl}^{-1}$$

$$\therefore \text{Concentration of } Na_2CO_3 = 2.0352\text{gl}^{-1}$$

### QUESTIONS FOR PRACTICE

- 48.0cm<sup>3</sup> of a solution containing 7.0g per litre of potassium hydroxide required 25.0cm<sup>3</sup> of 0.12M solution phosphoric acid for complete neutralization.
  - Calculate the molar ratio in which the phosphoric acid reacts with potassium hydroxide.
  - Write the equation for the above reaction.
- 1.65g of impure barium hydroxide was reacted with 100cm<sup>3</sup> of 0.2M hydrochloric acid. The excess acid required 10.0cm<sup>3</sup> of sodium hydroxide solution. If 25.0cm<sup>3</sup> of the sodium hydroxide solution required 28.5cm<sup>3</sup> of the hydrochloric acid for complete neutralization, calculate the percentage purity of barium hydroxide. ( $Ba = 137$ ;  $O = 16$ ;  $H = 1$ )
- A solution of 1.92g of potassium tetraoxalate  $H_2C_2O_4 \cdot KHC_2O_4 \cdot nH_2O$  required 22.7cm<sup>3</sup> of 1M solution sodium hydroxide. Determine the value of n ( $K = 39$ ;  $C = 12$ ;  $H = 1$ ;  $O = 16$ )
- 1.4g of sample of an iron wire was dissolved in dilute sulphuric acid and the solution made to 250cm<sup>3</sup> with distilled water. If 25cm<sup>3</sup> of this solution required 25.37cm<sup>3</sup> of a solution containing 2.33g of manganate(VII) ion per litre. Calculate the percentage of iron in the sample. ( $Mn = 55$ ;  $Fe = 56$ ;  $O = 16$ )
- 25.0cm<sup>3</sup> of a solution containing iron(II) and iron(III) sulphate required 18.50cm<sup>3</sup> of 0.1M potassium permanganate in acidic conditions. After complete reduction using zinc amalgam, 25.0cm<sup>3</sup> of the solution now required 32.6cm<sup>3</sup> of the same potassium permanganate solution. Calculate the concentration in grams per litre of each anhydrous sulphate.
- A solution Z containing iron(II) and iron(III) sulphate and potassium oxalate together with excess sulphuric acid was analyzed as follows;
  - 25.0cm<sup>3</sup> of Z was warmed and titrated and it required 30.0cm<sup>3</sup> of 0.02M potassium manganate(VII) solution.
  - 25.0cm<sup>3</sup> of Z was reduced, warmed and titrated. It required 40.0cm<sup>3</sup> of 0.02M potassium manganate(VII) solution.
  - 25.0cm<sup>3</sup> of Z was reduced, and required 28.0cm<sup>3</sup> of 0.02M potassium manganate(VII) solution.Calculate the concentration in grams per litre of iron(II) ions, iron(III) ions and the oxalate ion.
- 6.4g of copper(II) salt were dissolved in water and the solution made to 250cm<sup>3</sup>. To 25.0cm<sup>3</sup> of this solution was added excess potassium iodide solution and the iodine

- liberated required  $23.5\text{cm}^3$  of  $0.1\text{M}$  sodium thiosulphate solution. Calculate the percentage of copper in the salt. ( $Cu = 63.5$ )
8.  $1.25\text{g}$  of a chromium compound was converted to a chromate and the resultant solution made to  $1\text{dm}^3$ . To  $25.0\text{cm}^3$  of this solution was added  $25\text{cm}^3$  of  $2\text{M}$  sulphuric acid followed by  $25.0\text{cm}^3$  of  $0.1\text{M}$  potassium iodide solution. The liberated iodine required  $24.8\text{cm}^3$  of a solution containing  $12.5\text{g}$  per litre of potassium thiosulphate. Calculate the percentage of chromium in the compound.
  9.  $1.6\text{g}$  of a mixture of sodium sulphite,  $\text{Na}_2\text{SO}_3 \cdot 7\text{H}_2\text{O}$  and sodium thiosulphate,  $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$  was added to  $50\text{cm}^3$  of  $0.05\text{M}$  iodine solution. After complete reaction, the excess iodine was found to required  $24.95\text{cm}^3$  of  $0.05\text{M}$  sodium thiosulphate solution. Calculate the percentage of sodium sulphite crystals in the original mixture of salts.
  10.  $2.5\text{g}$  of a chlorate  $\text{MClO}_3$  ( $M$  representing the symbol of a metal) was heated strongly until no further change occurred. The residue formed was dissolved in water and the solution made up to  $250\text{cm}^3$ . If  $20.0\text{cm}^3$  of this solution required  $16.3\text{cm}^3$  of  $0.1\text{M}$  solution of silver nitrate, determine the atomic mass of  $M$ .
  11.  $1.64\text{g}$  of bromoalkane was heated with excess sodium hydroxide solution. The resultant solution was cooled and acidified with excess nitric acid and the diluted to  $100\text{cm}^3$ . If  $10.0\text{cm}^3$  of the acidified solution required  $13.0\text{cm}^3$  of  $0.1\text{M}$  silver nitrate solution for complete precipitation of silver bromide, Calculate the molar mass of the bromoalkane.
  12.  $25.0\text{cm}^3$  of solution **P** containing potassium chloride and potassium iodide required  $42.3\text{cm}^3$  of  $0.125\text{M}$  silver nitrate solution for complete precipitation.  $25\text{cm}^3$  of the same solution when treated with excess dilute hydrochloric acid and potassium iodate liberated iodine which was absorbed in  $50.0\text{cm}^3$  of  $0.1\text{M}$  sodium thiosulphate. Titration of the unused (excess) thiosulphate then required  $26.0\text{cm}^3$  of  $0.06\text{M}$  iodine. Calculate the concentration in grams per litre of the potassium chloride and potassium iodide in **P**.
  13.  $25.0\text{cm}^3$  of a solution containing  $13.54\text{g}$  of hydrated tin(II) chloride ( $\text{SnCl}_2 \cdot n\text{H}_2\text{O}$ ) per litre required  $30.0\text{cm}^3$  of  $0.05\text{M}$  solution of iodine. Calculate the percentage of water of crystallization in the tin(II) chloride salt.
  14.  $3.0\text{g}$  of sample of haematite,  $\text{Fe}_2\text{O}_3$ , were dissolved in dilute hydrochloric acid and the solution diluted to  $250\text{cm}^3$ .  $25.0\text{cm}^3$  of this solution after reduction with tin(II) chloride required  $26.6\text{cm}^3$  of  $0.02\text{M}$  potassium dichromate for oxidation. Calculate the percentage of iron(III) oxide in the ore.

15. 0.35g of compound sample containing a mixture of potassium chloride and potassium chlorate were dissolved in water and the resultant solution made up to 250cm<sup>3</sup>. To 25cm<sup>3</sup> of this solution was added to 20cm<sup>3</sup> of 2M hydrochloric acid followed by excess potassium iodide solution. The iodine liberated required 24.6cm<sup>3</sup> of 0.02M solution of sodium thiosulphate for complete reaction using starch.
16. 1.27g of iron(II) ethanedioate were dissolved and made up to 250cm<sup>3</sup>. 25.0cm<sup>3</sup> of this solution required 26.5cm<sup>3</sup> of acidified potassium manganate(VII) solution for complete oxidation.
- (a) Calculate the mole ratio of reaction between potassium manganate(VII) and iron(II) ethanedioate.
- (b) Write the molecular equation for the reaction.
17. 1.2g of an impure sample of calcium carbonate was dissolved in dilute hydrochloric acid and the calcium was precipitated as an oxalate by addition of a solution of ammonium oxalate. The precipitate was filtered, washed and dissolved in dilute sulphuric acid. This solution was then diluted to 1dm<sup>3</sup> with distilled water. 25.0cm<sup>3</sup> of the resultant solution was titrated with 0.002M potassium manganate(VII) solution and required 37.5cm<sup>3</sup>. Determine the percentage purity of calcium carbonate in the sample.
18. 1.43g of an impure salt of potassium permanganate was dissolved in water and the resultant solution made up to 250cm<sup>3</sup>. 25.0cm<sup>3</sup> of this solution was pipetted and acidified with sulphuric acid and warmed to about 70°C. The hot solution required 26.3cm<sup>3</sup> of a solution of anhydrous sodium ethanedioate made by dissolving 1.58g in 250cm<sup>3</sup> of solution. Determine the percentage purity of potassium permanganate.
19. 10cm<sup>3</sup> of a saturated solution of sulphur dioxide were added to 200cm<sup>3</sup> of 0.25M iodine solution and the resultant solution made to 1dm<sup>3</sup>. 25.0cm<sup>3</sup> of the solution required 16.7cm<sup>3</sup> of 0.1M sodium thiosulphate solution for complete reaction with excess iodine; calculate the solubility of sulphur dioxide in gdm<sup>-3</sup>.
20. Chlorine gas was bubbled through a solution containing 15.8g of sodium thiosulphate per litre. When the reaction was complete, the excess of dissolved chlorine was removed and the resulting solution gave the following data;
- (a) 20.0cm<sup>3</sup> required 20.0cm<sup>3</sup> of 1M sodium hydroxide for neutralization.
- (b) 20.0cm<sup>3</sup> required 16.0cm<sup>3</sup> of 1M silver nitrate solution for complete precipitation.
- (c) 20.0cm<sup>3</sup> gave 0.932g of barium sulphate when excess barium chloride solution was added.
- Deduce the equation of reaction between chlorine and sodium thiosulphate.

21. 25.0cm<sup>3</sup> of a solution anhydrous sodium sulphite were added to 50cm<sup>3</sup> of 0.05M iodine solution acidified with hydrochloric acid. The excess iodine required 17.5cm<sup>3</sup> of 0.05M sodium thiosulphate. Calculate the mass of sodium sulphite in 1 litre of solution.
22. 1.30g of a sample of sodium nitrite contaminated with traces of sodium nitrate was dissolved in distilled water and the resultant solution made up to 250cm<sup>3</sup>. 25cm<sup>3</sup> of this solution required 28.6cm<sup>3</sup> of 0.02M potassium dichromate. Calculate the percentage purity of sodium nitrite salt.
23. A compound **Q** contains carbon, hydrogen, nitrogen and oxygen. When 0.225g of **Q** was heated with sodium hydroxide solution and the ammonia evolved was passed into 25.0cm<sup>3</sup> of 0.1M sulphuric acid. The excess acid required 19.1cm<sup>3</sup> of 0.1M sodium hydroxide for complete neutralization. 0.195g of sample **Q** gave on complete oxidation 0.325g of carbon dioxide and 0.168g of water. A solution of 9.12g of **Q** in 500cm<sup>3</sup> of water froze at -0.465°C. if the cryoscopic constant for water is 1.86kgmol<sup>-1</sup>
- (a) Determine the molecular formula of **Q**.  
 (b) Identify **Q**.
24. Ammonium ions react with methanal to produce a neutral organic compound and hydrogen ions according to the following equation;
- $$4NH_4^+(aq) + 6HCHO \rightarrow C_6H_{12}N_4 + 4H^+ + 6H_2O$$
- 5.0g of the fertilizer containing ammonium sulphate was dissolved in water and the solution made up to 500cm<sup>3</sup>. 25cm<sup>3</sup> portions of this solution were treated with aqueous methanal and allowed to stand for a few minutes. The liberated acid required 22.40cm<sup>3</sup> of 0.1M sodium hydroxide solution for complete neutralization. Calculate the percentage by mass of ammonium sulphate in the fertilizer.