CHEMISTRY NOTES

FORM 4

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**(a)Gas laws**

1. Matter is made up of small particle in accordance to Kinetic Theory of matter:

Naturally, there are basically **three** states of matter: **Solid, Liquid** and **gas:**

(i)A solid is made up of particles which are very closely packed with a definite/fixed shape and fixed/definite volume /occupies definite space. It has a very high density.

(ii) A liquid is made up of particles which have some degree of freedom. It thus has no definite/fixed shape. It takes the shape of the container it is put. A liquid has fixed/definite volume/occupies definite space.

(iii)A gas is made up of particles free from each other. It thus has no definite /fixed shape. It takes the shape of the container it is put. It has no fixed/definite volume/occupies every space in a container.

2.Gases are affected by **physical conditions**. There are **two** physical conditions:

(i)**Temperature**

(ii)**Pressure**

3. The SI unit of temperature is **Kelvin**(**K**).

**Degrees** Celsius/Centigrade(**oC**) are also used.

The two units can be interconverted from the relationship:

**oC + 273= K**

**K -273 = oC**

Practice examples

1. Convert the following into Kelvin.

(i) O **oC**

**oC + 273 = K substituting :** O oC + 273 = **273 K**

(ii) -273 **oC**

**oC + 273 = K substituting :** -273oC + 273 = **0 K**

(iii) 25 **oC**

**oC + 273 = K substituting :** 25 oC + 273 = **298 K**

(iv) 100 **oC**

**oC + 273 = K substituting :** 100 oC + 273 = **373 K**

2. Convert the following into degrees Celsius/Centigrade(oC).

(i) 10 K

**K -273 = oC substituting:** 10 – 273 = **-263 oC**

(ii) (i) 1 K

**K -273 = oC substituting:** 1 – 273 =  **-272 oC**

(iii) 110 K

**K -273 = oC substituting: 110 – 273 = -163 oC**

(iv) -24 K

**K -273 = oC substituting: -24 – 273 = -297 oC**

The **standard** temperature is **273K** = **0 oC**.

The **room** temperature is assumed to be **298K** = **25oC**

4. The SI unit of pressure is Pascal(**Pa**) / Newton per metre squared (**Nm-2**) . Millimeters’ of mercury(**mmHg**) ,centimeters of mercury(**cmHg**) and **atmospheres** are also commonly used.

The units are **not** interconvertible but Pascals(Pa) are equal to Newton per metre squared(Nm-2).

The **standard** pressure is the **atmospheric** pressure.

Atmospheric pressure is **equal** to about:

(i)101325 Pa

(ii)101325 Nm-2

(iii)760 mmHg

(iv)76 cmHg

(v)one atmosphere.

5. Molecules of gases are always in continuous random motion at high speed. This motion is affected by the physical conditions of temperature and pressure.

Physical conditions change the volume occupied by gases in a **closed** system.

The effect of physical conditions of temperature and pressure was investigated and expressed in both Boyles and Charles laws.

6. Boyles law states that

“**the volume of a fixed mass of a gas is inversely proportional to the pressure at constant/fixed temperature** ”

Mathematically:

Volume α 1 **(Fixed /constant Temperature)**

Pressure

V α 1 **(Fixed /constant T)** ie **PV = Constant**(k)

P

From Boyles law , an **increase** in pressure of a gas cause a **decrease** in volume. i.e **doubling** the pressure cause the volume to be **halved**.

Graphically therefore a plot of volume(**V**) **against** pressure (**P**) produces a **curve**.

V

P

Graphically a plot of volume(**V**) **against** inverse/reciprocal of pressure (**1/p**) produces a **straight line**

**V**

**1/P**

For **two** gases then **P1 V1 = P2 V2**

P1 = Pressure of gas 1

V1 = Volume of gas 1

P2 = Pressure of gas 2

V2 = Volume of gas 2

Practice examples:

**1. A fixed mass of gas at 102300Pa pressure has a volume of 25cm3.Calculate its volume if the pressure is doubled.**

Working

P1 V1 = P2 V2 Substituting :102300 x 25 = (102300 x **2**) x V2

V2 = 102300 x 25 = **12.5cm3**

(102300 x 2)

**2. Calculate the pressure which must be applied to a fixed mass of 100cm3 of Oxygen for its volume to triple at 100000Nm-2.**

P1 V1 = P2 V2 Substituting :100000 x 100 = P2 x (100 x **3**)

V2 = 100000 x 100 = **33333.3333 Nm-2**

(100 x 3)

**3.A 60cm3 weather ballon full of Hydrogen at atmospheric pressure of 101325Pa was released into the atmosphere. Will the ballon reach stratosphere where the pressure is 90000Pa?**

P1 V1 = P2 V2 Substituting :101325 x 60 = 90000 x V2

V2 = 101325 x 60 = **67.55 cm3**

90000

The new volume at 67.55 cm3 **exceed** ballon capacity of 60.00 cm3.It will **burst** before reaching destination.

7.Charles law states that“**the volume of a fixed mass of a gas is directly proportional to the absolute temperature at constant/fixed pressure** ”

Mathematically:

Volume α Pressure **(Fixed /constant pressure)**

V α T  **(Fixed /constant P)** ie **V = Constant**(k)

T

From Charles law , an **increase** in temperature of a gas cause an **increase** in volume. i.e **doubling** the temperature cause the volume to be **doubled**.

Gases expand/increase by **1/273** by volume on heating.Gases contact/decrease by **1/273** by volume on cooling at constant/fixed pressure.

The volume of a gas continue decreasing with decrease in temperature until at **-273oC /0 K** the volume is **zero**. i.e. there is no gas.

This temperature is called **absolute zero.** It is the **lowest** temperature at which a gas **can** exist.

Graphically therefore a plot of volume(**V**) **against** Temperature(**T**) in:

(**i**)oC produces a **straight line** that is **extrapolated** to the absolute zero of -273**oC** .

V

**-273oC** **0oC**

T(**oC**)

(**ii**)Kelvin/K produces a **straight line** from absolute zero of **O Kelvin**

**V**

**0 T(Kelvin)**

For **two** gases then **V1  = V2**

**T1  T2**

T1 = Temperature **in Kelvin** of gas 1

V1 = Volume of gas 1

T2 = Temperature **in Kelvin** of gas 2

V2 = Volume of gas 2

Practice examples:

**1. 500cm3 of carbon(IV)oxide at 0oC was transfered into a cylinder at -4oC. If the capacity of the cylinder is 450 cm3,explain what happened.**

V1  = V2 substituting 500 = V2

T1  T2 (0 +273) (-4 +273)

= 500 x (-4 x 273) = **492.674cm3**

(0 + 273)

The capacity of cylinder (500cm3) is **less** than new volume(492.674cm3).

**7.326cm3**(500-492.674cm3)of carbon(IV)oxide gas did not fit into the cylinder.

**2. A mechanic was filling a deflated tyre with air in his closed garage using a hand pump. The capacity of the tyre was 40,000cm3 at room temperature. He rolled the tyre into the car outside. The temperature outside was 30oC.Explain what happens.**

V1  = V2 substituting 40000 = V2

T1  T2 (**25** +273) (**30** +273)

= 40000 x (30 x 273) = **40671.1409cm3**

(25 + 273)

The capacity of a tyre (40000cm3) is **less** than new volume(40671.1409cm3).

The tyre thus bursts.

**3. A hydrogen gas balloon with 80cm3 was released from a research station at room temperature. If the temperature of the highest point it rose is -30oC , explain what happened.**

V1  = V2 substituting 80 = V2

T1  T2 (**25** +273) (-**30** +273)

= 80 x (-30 x 273) = **65.2349cm3**

(25 + 273)

The capacity of balloon (80cm3) is **more** than new volume (65.2349cm3).

The balloon thus remained intact.

8. The continuous random motion of gases differ from gas to the other.The movement of molecules (of a gas) from region of high concentration to a region of low concentration is called **diffusion.**

The rate of diffusion of a gas depends on its density. i.e. **The higher the rate of diffusion, the less dense the gas**.

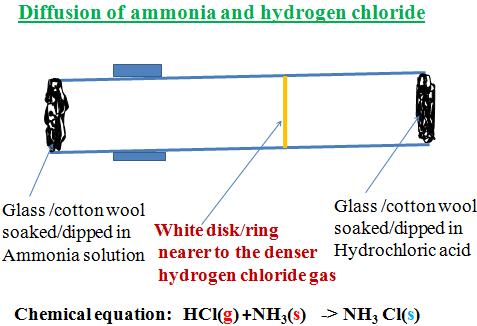
The density of a gas depends on its molar mass/relative molecular mass. i.e. **The** higher **the density the** higher **the molar mass/relative atomic mass and thus the** lower **the rate of diffusion.**

Examples

1.Carbon (IV)oxide(CO2) has a molar mass of 44g.Nitrogen(N2)has a molar mass of 28g. (N2)is thus lighter/less dense than Carbon (IV)oxide(CO2). N2 diffuses faster than CO2.

2.Ammonia(NH3) has a molar mass of 17g.Nitrogen(N2)has a molar mass of 28g. (N2)is thus about **twice** lighter/less dense than Ammonia(NH3). Ammonia(NH3)diffuses twice faster than N2.

3. Ammonia(NH3) has a molar mass of 17g.Hydrogen chloride gas has a molar mass of 36.5g.Both gases on contact react to form **white fumes** of ammonium chloride .When a glass/cotton wool dipped in ammonia and another glass/cotton wool dipped in hydrochloric acid are placed at opposite ends of a glass tube, both gases diffuse towards each other. A white disk appears near to glass/cotton wool dipped in hydrochloric acid. This is because hydrogen chloride is heavier/denser than Ammonia and thus its rate of diffusion is lower .



The rate of diffusion of a gas is in accordance to **Grahams law of diffusion**.Grahams law states that:

**“the rate of diffusion of a gas is inversely proportional to the square root of its density, at the same/constant/fixed temperature and pressure”**

Mathematically

**R α 1**  and since density is proportional to mass then **R α 1**

**√ p √ m**

For two gases then:

**R1  = R2** where: R**1** and R**2** is the **rate** of diffusion of **1**st and **2**nd gas.

**√M2 √M1** M**1** and M**2** is the **molar mass** of **1**st and **2**nd gas.

Since rate is inverse of time. i.e. the higher the rate the less the time:

For two gases then:

T**1**= T2 where: T**1** and T**2** is the **time taken** for **1**st and **2**nd gas to diffuse.

**√**M1 **√**M2  M**1** and M**2** is the **molar mass** of **1**st and **2**nd gas.

**Practice examples**:

**1. It takes 30 seconds for 100cm3 of carbon(IV)oxide to diffuse across a porous plate. How long will it take 150cm3 of nitrogen(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,N=14.0=16.0)**

Molar mass CO2=44.0 Molar mass NO2=46.0

Method 1

100cm3 CO2  takes 30seconds

150cm3 takes 150 x30 = 45seconds

100

T CO2 = √ molar mass CO2 => 45seconds = √ 44.0

T NO2  √ molar mass NO2 T NO2 √ 46.0

T NO2 =45seconds x √ 46.0 = **46.0114** seconds

√ 44.0

Method 2

100cm3 CO2  takes 30seconds

1cm3 takes 100 x1 = **3.3333cm3sec-1**

30

R CO2 = √ molar mass NO2 => 3.3333cm3sec-1 = √ 46.0

R NO2  √ molar mass CO2 R NO2 √ 44.0

R NO2 = 3.3333cm3sec-1 x √ 44.0 = **3.2601**cm3sec-1

√ 46.0

3.2601cm3 takes 1seconds

150cm3 take 150cm3 = **46.0109seconds**

3.2601cm3

**2. How long would 200cm3 of Hydrogen chloride take to diffuse through a porous plug if carbon(IV)oxide takes 200seconds to diffuse through.**

Molar mass CO2 = 44g Molar mass HCl = 36.5g

T CO2 = √ molar mass CO2 => 200 seconds = √ 44.0

T HCl √ molar mass HClT HCl√ 36.5

T HCl = 200seconds x √ 36.5 = **182.1588** seconds

√ 44.0

**3. Oxygen gas takes 250 seconds to diffuse through a porous diaphragm. Calculate the molar mass of gas Z which takes 227 second to diffuse.**

Molar mass O2 = 32g Molar mass Z = x g

T O2 = √ molar mass O2 => 250 seconds = √ 32.0

T Z √ molar mass Z227seconds√ x

√ x = 227seconds x √ 32 = **26.3828** grams

250

**4. 25cm3 of carbon(II)oxide diffuses across a porous plate in 25seconds. How long will it take 75cm3 of Carbon(IV)oxide to diffuse across the same plate under the same conditions of temperature and pressure. (C=12.0,0=16.0)**

Molar mass CO2 = 44.0 Molar mass CO = 28.0

Method 1

25cm3 CO takes 25seconds

75cm3 takes 75 x25 = 75seconds

25

T CO2 = √ molar mass CO2 => T CO2seconds = √ 44.0

T CO √ molar mass CO75√ 28.0

T CO2 =75seconds x √ 44.0 = **94.0175** seconds

√ 28.0

Method 2

25cm3 CO2  takes 25seconds

1cm3 takes 25 x1 = **1.0cm3sec-1**

25

R CO2 = √ molar mass CO => x cm3sec-1 = √ 28.0

R CO √ molar mass CO2 1.0cm3sec-1√ 44.0

R CO2 = 1.0cm3sec-1 x √ 28.0 = **0.7977**cm3sec-1

√ 44.0

0.7977cm3 takes 1 seconds

75cm3 takes 75cm3 = **94.0203seconds**

0.7977cm3

**(b)Introduction to the mole, molar masses and Relative atomic masses**

1. The mole is the **SI** unit of the **amount** of substance.

2. The number of particles e.g. atoms, ions, molecules, electrons, cows, cars are all measured in terms of moles.

3. The number of particles in one mole is called the **Avogadros Constant**. It is denoted “**L**”.

The Avogadros Constant contain **6.023 x10 23** particles. i.e.

1mole = 6.023 x10 23 particles = 6.023 x10 **23**

2 moles = 2 x 6.023 x10 23 particles = 1.205 x10 **24**

0.2 moles = 0.2 x 6.023 x10 23 particles = 1.205 x10 **22**

0.0065 moles = 0.0065 x 6.023 x10 23 particles = 3.914 x10 **21**

3. The mass of one mole of a substance is called **molar mass**. The molar mass of:

(i)an **element** has mass equal to relative **atomic** mass /RAM(in grams)of the element e.g.

Molar mass of carbon(C)= relative atomic mass = 12.0g

6.023 x10 23 particles of carbon = 1 mole =12.0 g

Molar mass of sodium(Na) = relative atomic mass = 23.0g

6.023 x10 23 particles of sodium = 1 mole =23.0 g

Molar mass of Iron (Fe) = relative atomic mass = 56.0g

6.023 x10 23 particles of iron = 1 mole =56.0 g

(ii)a **molecule** has mass equal to relative molecular mass /RMM (in grams)of the molecule. Relative molecular mass is the **sum** of the relative atomic masses of the elements making the molecule.

The number of atoms making a molecule is called **atomicity**. Most **gaseous** molecules are **diatomic (e.g. O2**, **H2**, **N2**, **F2**, **Cl2**, **Br2**, **I2**)noble gases are **monoatomic(e.g. He, Ar, Ne, Xe)**,Ozone gas(**O3**) is **triatomic** e.g.

Molar mass **Oxygen molecule(O2)** =relative molecular mass =(16.0x 2)g =32.0g

6.023 x10 23 particles of Oxygen molecule = 1 mole = 32.0 g

Molar mass **chlorine molecule(Cl2)** =relative molecular mass =(35.5x 2)g =71.0g

6.023 x10 23 particles of chlorine molecule = 1 mole = 71.0 g

Molar mass **Nitrogen molecule(N2)** =relative molecular mass =(14.0x 2)g =28.0g

6.023 x10 23 particles of Nitrogen molecule = 1 mole = 28.0 g

(ii)a **compound** has mass equal to relative formular mass /RFM (in grams)of the molecule. Relative formular mass is the **sum** of the relative atomic masses of the elements making the compound. e.g.

(i)Molar mass **Water(H2O)** = relative formular mass =[(1.0 x 2 ) + 16.0]g =18.0g

6.023 x10 23 particles of Water molecule = 1 mole = 18.0 g

6.023 x10 23 particles of Water molecule has:

- **2** x 6.023 x10 23 particles of Hydrogen atoms

**-1** x 6.023 x10 23 particles of Oxygen atoms

(ii)Molar mass **sulphuric(VI)acid(H2SO4)** = relative formular mass

=[(1.0 x 2 ) + 32.0 + (16.0 x 4)]g =98.0g

6.023 x10 23 particles of sulphuric(VI)acid(H2SO4) = 1 mole = 98.0g

6.023 x10 23 particles of sulphuric(VI)acid(H2SO4) has:

- **2** x 6.023 x10 23 particles of **H**ydrogen atoms

-**1** x 6.023 x10 23 particles of **S**ulphur atoms

**-4** x 6.023 x10 23 particles of **O**xygen atoms

(iii)Molar mass **sodium carbonate(IV)(Na2CO3)** = relative formular mass

=[(23.0 x 2 ) + 12.0 + (16.0 x 3)]g =106.0g

6.023 x10 23 particles of sodium carbonate(IV)(Na2CO3) = 1 mole = 106.0g

6.023 x10 23 particles of sodium carbonate(IV)(Na2CO3) has:

- **2** x 6.023 x10 23 particles of **S**odium atoms

-**1** x 6.023 x10 23 particles of **C**arbon atoms

**-3** x 6.023 x10 23 particles of **O**xygen atoms

(iv)Molar mass **Calcium carbonate(IV)(CaCO3)** = relative formular mass

=[(40.0+ 12.0 + (16.0 x 3)]g =100.0g.

6.023 x10 23 particles of Calcium carbonate(IV)(CaCO3) = 1 mole = 100.0g

6.023 x10 23 particles of Calcium carbonate(IV)(CaCO3) has:

- **1** x 6.023 x10 23 particles of **C**alcium atoms

-**1** x 6.023 x10 23 particles of **C**arbon atoms

**-3** x 6.023 x10 23 particles of **O**xygen atoms

(v)Molar mass **Water(H2O)** = relative formular mass

=[(2 x 1.0 )+ 16.0 ]g =18.0g

6.023 x10 23 particles of Water(H2O) = 1 mole = 18.0g

6.023 x10 23 particles of Water(H2O) has:

- **2** x 6.023 x10 23 particles of Hydrogen atoms

**-2** x 6.023 x10 23 particles of **O**xygen atoms

**Practice**

1. Calculate the number of moles present in:

(i)0.23 g of Sodium atoms

Molar mass of Sodium atoms = 23g

Moles = mass in grams = > 0.23g = **0.01moles**

Molar mass 23

(ii) 0.23 g of Chlorine atoms

Molar mass of Chlorine atoms = 35.5 g

Moles = mass in grams = > 0.23g = **0.0065moles /6.5 x 10-3 moles**

Molar mass 35.5

(iii) 0.23 g of Chlorine molecules

Molar mass of Chlorine molecules =( 35.5 x 2) = **71.0** g

Moles = mass in grams = > 0.23g = **0.0032moles /3.2 x 10-3 moles**

Molar mass 71

(iv) 0.23 g of dilute sulphuric(VI)acid

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams = > 0.23g = **0.0023moles /2.3 x 10-3 moles**

Molar mass 98

2. Calculate the number of atoms present in:(Avogadros constant L = 6.0 x 10 23)

(i) 0.23 g of dilute sulphuric (VI)acid

**Method I**

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

Moles = mass in grams = > 0.23g = **0.0023moles /2.3 x 10-3 moles**

Molar mass 98

1 mole has 6.0 x 10 23 atoms

2.3 x 10-3 moles has (2.3 x 10-3 x 6.0 x 10 23) = **1.38 x 10 21**atoms

1

**Method II**

Molar mass of H2SO4 = [(2 x 1) + 32 + (4 x14)] = **98.0**g

**98.0**g = 1 mole has 6.0 x 10 23 atoms

0.23 g therefore has (0.23 g x 6.0 x 10 23 )= **1.38 x 10 21**atoms

98

(ii)0.23 g of sodium carbonate(IV)decahydrate

Molar mass of Na2CO3.10H2 O=

[(2 x 23) + 12 + (3 x16) + (10 x 1.0) + (10 x 16)] = **276.0**g

**Method I**

Moles = mass in grams = > 0.23g = **0.00083moles /**

Molar mass 276 **8.3 x 10-4 moles**

1 mole has 6.0 x 10 23 atoms

8.3 x 10-4 moles has (8.3 x 10-4 moles x 6.0 x 10 23) = **4.98 x 10 20**atoms

1

**Method II**

276.0g = 1 mole has 6.0 x 10 23 atoms

0.23 g therefore has (0.23 g x 6.0 x 10 23 )= **4.98 x 10 20**atoms

276.0

(iii)0.23 g of Oxygen gas

Molar mass of O2 = (2 x16) = **32.0** g

**Method I**

Moles = mass in grams = > 0.23g = **0.00718moles /**

Molar mass 32 **7.18 x 10-3 moles**

1 mole has **2** x 6.0 x 10 23 **atoms** in O2

7.18 x 10-3moles has (7.18 x 10-3moles x **2** x 6.0 x 10 23) =**8.616 x 10 21**atoms

1

**Method II**

32.0g = 1 mole has **2** x 6.0 x 10 23 **atoms** in O2

0.23 g therefore has (0.23 g x **2** x 6.0 x 10 23 )= **8.616 x 10 21**atoms

32.0

(iv)0.23 g of Carbon(IV)oxide gas

Molar mass of CO2 = [12 + (2 x16)] = **44.0** g

**Method I**

Moles = mass in grams = > 0.23g = **0.00522moles /**

Molar mass 44 **5.22 x 10-3 moles**

1 mole has **3** x 6.0 x 10 23 **atoms** in CO2

7.18 x 10-3moles has (5.22 x 10-3moles x **3** x 6.0 x 10 23) =**9.396 x 10 21**atoms

1

**Method II**

44.0g = 1 mole has **3** x 6.0 x 10 23 **atoms** in CO2

0.23 g therefore has (0.23 g x 3 x 6.0 x 10 23 )= **9.409 x 10 21**atoms

44.0

**(c)Empirical and molecular formula**

1.The empirical formula of a compound is its simplest formula. It is the simplest whole number ratios in which atoms of elements combine to form the compound. 2.It is mathematically the lowest common multiple (LCM) of the atoms of the elements in the compound

3.Practically the empirical formula of a compound can be determined as in the following examples.

**To determine the empirical formula of copper oxide**

**(a)**Method 1:From copper to copper(II)oxide

**Procedure.**

Weigh a clean dry covered crucible(**M1**).Put two spatula full of copper **powder** into the crucible. Weigh again (**M2**).Heat the crucible on a strong Bunsen flame for five minutes. Lift the lid, and swirl the crucible carefully using a pair of tong. Cover the crucible and continue heating for another five minutes. Remove the lid and stop heating. Allow the crucible to cool. When cool replace the lid and weigh the contents again (**M3**).

**Sample results**

|  |  |
| --- | --- |
| Mass of crucible(**M1**) | 15.6g |
| Mass of crucible + copper before heating(**M2**) | 18.4 |
| Mass of crucible + copper after heating(**M3**) | 19.1 |

**Sample questions**

**1. Calculate the mass of copper powder used.**

Mass of crucible + copper before heating(**M2**) = 18.4

Less Mass of crucible(**M1**) = - 15.6g

Mass of copper **2.8 g**

**2. Calculate the mass of Oxygen used to react with copper.**

Method I

Mass of crucible + copper after heating(**M3**) = 19.1g

Mass of crucible + copper before heating(**M2**) = - 18.4g

Mass of Oxygen =  **0.7 g**

Method II

Mass of crucible + copper after heating(**M3**) = 19.1g

Mass of crucible = - 15.6g

Mass of copper(II)Oxide = 3.5 g

Mass of copper(II)Oxide = 3.5 g

Mass of copper = - 2.8 g

Mass of Oxygen =  **0.7 g**

**3. Calculate the number of moles of:**

(i) copper used (Cu = 63.5)

number of moles of copper = mass used => 2.8 = **0.0441moles**

Molar mass 63.5

(ii) Oxygen used (O = 16.0)

number of moles of oxygen = mass used => 0.7 = **0.0441moles**

Molar mass 16.0

**4. Determine the mole ratio of the reactants**

Moles of copper = 0.0441moles = 1 => Mole ratio Cu: O = **1:1**

Moles of oxygen 0.0441moles 1

**5.What is the empirical, formula of copper oxide formed.**

CuO (copper(II)oxide

**6. State and explain the observations made during the experiment.**

Observation

Colour change from **brown** to **black**

Explanation

Copper powder is brown. On heating it reacts with oxygen from the air to form black copper(II)oxide

**7. Explain why magnesium ribbon/shavings would be unsuitable in a similar experiment as the one above.**

Hot magnesium generates enough heat energy to react with **both** Oxygen and Nitrogen in the air forming a white solid mixture of Magnesuin oxide and magnesium nitride. This causes experimental mass errors.

**(b)**Method 2:From copper(II)oxide to copper

**Procedure.**

Weigh a clean dry porcelain boat (**M1**). Put two spatula full of copper(II)oxide **powder** into the crucible. Reweigh the porcelain boat (**M2**).Put the porcelain boat in a glass tube and set up the apparatus as below;



Pass slowly(to prevent copper(II)oxide from being blown away)a stream of either dry Hydrogen /ammonia/laboratory gas/ carbon(II)oxide gas for about two minutes from a suitable generator.

When all the in the apparatus set up is driven out ,heat the copper(II)oxide strongly for about five minutes until there is no further change. Stop heating.

Continue passing the gases until the glass tube is cool.

Turn off the gas generator.

Carefully remove the porcelain boat form the combustion tube.

Reweigh (**M3**).

**Sample results**

|  |  |
| --- | --- |
| Mass of boat(**M1**) | 15.6g |
| Mass of boat before heating(**M2**) | 19.1 |
| Mass of boat after heating(**M3**) | 18.4 |

**Sample questions**

**1. Calculate the mass of copper(II)oxide used.**

Mass of boat before heating(**M2**) = 19.1

Mass of empty boat(**M1**) = - 15.6g

Mass of copper(II)Oxide **3.5 g**

**2. Calculate the mass of**

**(i) Oxygen.**

Mass of boat before heating(**M2**) = 19.1

Mass of boat after heating (**M3**) = - 18.4g

Mass of oxygen = **0.7 g**

**(ii)Copper**

Mass of copper(II)Oxide = 3.5 g

Mass of oxygen = 0.7 g

Mass of oxygen = **2.8 g**

**3. Calculate the number of moles of:**

**(i) Copper used (Cu = 63.5)**

number of moles of copper = mass used => 2.8 = **0.0441moles**

Molar mass 63.5

**(ii) Oxygen used (O = 16.0)**

number of moles of oxygen = mass used => 0.7 = **0.0441moles**

Molar mass 16.0

**4. Determine the mole ratio of the reactants**

Moles of copper = 0.0441moles = 1 => Mole ratio Cu: O = **1:1**

Moles of oxygen 0.0441moles 1

**5.What is the empirical, formula of copper oxide formed.**

CuO (copper(II)oxide

**6. State and explain the observations made during the experiment.**

Observation

Colour change from **black** to **brown**

Explanation

Copper(II)oxide powder is black. On heating it is reduced by a suitable reducing agent to brown copper metal.

**7. Explain why magnesium oxide would be unsuitable in a similar experiment as the one above.**

Magnesium is high in the reactivity series. None of the above reducing agents is strong enough to reduce the oxide to the metal.

**8. Write the equation for the reaction that would take place when the reducing agent is:**

**(i) Hydrogen**

CuO(s) + H2(g) -> Cu(s) + H2O(l)

(Black) (brown) (colourless liquid form

on cooler parts )

**(ii)Carbon(II)oxide**

CuO(s) + CO (g) -> Cu(s) + CO2(g)

(Black) (brown) (colourless gas, form

white ppt with lime water )

**(iii)Ammonia**

3CuO(s) + 2NH3(g) -> 3Cu(s) + N2 (g) + 3H2O(l)

(Black) (brown) (colourless liquid form

on cooler parts )

**9. Explain why the following is necessary during the above experiment;**

**(i)A stream of dry hydrogen gas should be passed before heating copper (II) Oxide.**

Air combine with hydrogen in presence of heat causing an explosion

**(ii)A stream of dry hydrogen gas should be passed after heating copper (II) Oxide has been stopped.**

Hot metallic copper can be re-oxidized back to copper(II)oxide

**(iii) A stream of excess carbon (II)oxide gas should be ignited to burn**

Carbon (II)oxide is highly poisonous/toxic. On ignition it burns to form less toxic carbon (IV)oxide gas.

**10. State two sources of error in this experiment.**

(i)All copper(II)oxide may not be reduced to copper.

(ii)Some copper(II)oxide may be blown out the boat by the reducing agent.

4.Theoreticaly the empirical formula of a compound can be determined as in the following examples.

**(a)A oxide of copper contain 80% by mass of copper. Determine its empirical formula. (Cu = 63.5, 16.0)**

% of Oxygen = 100% - % of Copper => 100- 80 = **20%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Copper | Oxygen |
| Symbol | Cu | O |
| Moles present = % composition  Molar mass | 80  63.5 | 20  16 |
| Divide by the smallest value | 1.25  1.25 | 1.25  1.25 |
| Mole ratios | 1 | 1 |

Empirical formula is **CuO**

**(b)1.60g of an oxide of Magnesium contain 0.84g by mass of Magnesium. Determine its empirical formula(Mg = 24.0, 16.0)**

**Mass** of Oxygen = 1.60 – 0.84 => **0.56 g** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Magnesium | Oxygen |
| Symbol | Mg | O |
| Moles present = % composition  Molar mass | 0.84  24 | 0.56  16 |
| Divide by the smallest value | 0.35  0.35 | 0.35  0.35 |
| Mole ratios | 1 | 1 |

Empirical formula is **MgO**

**(c)An oxide of Silicon contain 47% by mass of Silicon. What is its empirical formula(Si = 28.0, 16.0)**

**Mass** of Oxygen = 100 – 47 => **53%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Silicon | Oxygen |
| Symbol | Si | O |
| Moles present = % composition  Molar mass | 47  28 | 53  16 |
| Divide by the smallest value | 1.68  1.68 | 3.31  1.68 |
| Mole ratios | 1 | 1.94 = 2 |

Empirical formula is **SiO2**

**(d)A compound contain 70% by mass of Iron and 30% Oxygen. What is its empirical formula(Fe = 56.0, 16.0)**

**Mass** of Oxygen = 100 – 47 => **53%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Silicon | Oxygen |
| Symbol | Si | O |
| Moles present = % composition  Molar mass | 47  28 | 53  16 |
| Divide by the smallest value | 1.68  1.68 | 3.31  1.68 |
| Mole ratios | 1 | 1.94 = 2 |

Empirical formula is **SiO2**

**2.During heating of a hydrated copper (II)sulphate(VI) crystals, the following readings were obtained:**

**Mass of evaporating dish =300.0g**

**Mass of evaporating dish + hydrated salt = 305.0g**

**Mass of evaporating dish + anhydrous salt = 303.2g**

**Calculate the number of water of crystallization molecules in hydrated copper (II)sulphate(VI)**

**(Cu =64.5, S = 32.0,O=16.0, H = 1.0)**

**Working**

Mass of Hydrated salt = 305.0g -300.0g = **5.0g**

Mass of anhydrous salt = 303.2 g -300.0g = **3.2 g**

Mass of water in hydrated salt = 5.0g -3.2 g = **1.8g**

Molar mass of water(H2O) = **18.0g**

Molar mass of anhydrous copper (II)sulphate(VI) (CuSO4) = **160.5g**

|  |  |  |
| --- | --- | --- |
| Element/compound | anhydrous copper (II)  sulphate(VI) | Oxygen |
| Symbol | Si | O |
| Moles present = composition by mass  Molar mass | 3,2  160.5 | 1.8  18 |
| Divide by the smallest value | 0.0199  0.0199 | 0.1  18 |
| Mole ratios | 1 | 5 |

The **empirical formula** of hydrated salt = CuSO4.**5**H2O

Hydrated salt has **five/5 molecules** of water of crystallizations

4. The molecular formula is the actual number of each kind of atoms present in a **molecule** of a compound.

The empirical formula of an ionic compound is the same as the chemical formula but for simple molecular structured compounds, the empirical formula may not be the same as the chemical formula.

The molecular formula is a multiple of empirical formula .It is determined from the relationship:

(i) **n**  = Relative formular mass

Relative empirical formula

where **n** is a whole number.

(ii) Relative empirical formula x **n** = Relative formular mass

where **n** is a whole number.

**Practice sample examples**

**1. A hydrocarbon was found to contain 92.3% carbon and the remaining Hydrogen.**

**If the molecular mass of the compound is 78, determine the molecular formula(C=12.0, H =1.0)**

**Mass** of Hydrogen = 100 – 92.3 => **7.7%** of Oxygen

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = % composition  Molar mass | 92.3  12 | 7.7  1 |
| Divide by the smallest value | 7.7  7.7 | 7.7  7.7 |
| Mole ratios | 1 | 1 |

Empirical formula is **CH**

The molecular formular is thus determined :

**n**  = Relative formular mass = 78 = 6

Relative empirical formula 13

The molecular formula is (C H ) x 6 = **C6H6**

**2. A compound of carbon, hydrogen and oxygen contain 54.55% carbon, 9.09% and remaining 36.36% oxygen.**

**If its relative molecular mass is 88, determine its molecular formula(C=12.0, H =1.0, O= 16.0)**

|  |  |  |  |
| --- | --- | --- | --- |
| Element | Carbon | Hydrogen | Oxygen |
| Symbol | C | H | O |
| Moles present = % composition  Molar mass | 54.55  12 | 9.09  1 | 36.36  16 |
| Divide by the smallest value | 4.5458  2.2725 | 9.09  2.2725 | 2.2725  2.2725 |
| Mole ratios | 2 | 4 | 1 |

Empirical formula is **C2H4O**

The molecular formula is thus determined :

**n**  = Relative formular mass = 88 = 2

Relative empirical formula 44

The molecular formula is (C2H4O ) x 2 = **C4H8O2**.

**4.A hydrocarbon burns completely in excess air to form 5.28 g of carbon (IV) oxide and 2,16g of water.**

**If the molecular mass of the hydrocarbon is 84, draw and name its molecular structure.**

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

Mass of carbon in CO2 = Mass of C in CO2 x mass of CO2 =>

Molar mass of CO2

12 x 5.28 = **1.44g√**

44

Mass of Hydrogen in H2O = Mass of C in H2Ox mass of H2O=>

Molar mass of H2O

2 x 2.16 = **0.24g√**

18

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = mass  Molar mass | 1.44g  12 | 0.24g**√**  1 |
| Divide by the smallest value | 0.12  0.12 | 0.24  0.12 |
| Mole ratios | 1 | 2**√** |

Empirical formula is **CH2√**

The molecular formular is thus determined :

**n**  = Relative formular mass = 84 = **6√**

Relative empirical formula 14

The molecular formula is (CH2 ) x 6 = **C6H12**. **√**

molecular name **Hexene√/Hex-1-ene (or any position isomer of Hexene)**

Molecular structure

**H H H H H H**

**H C C C C C C H√**

**H H H H**

**5. Compound A contain 5.2% by mass of Nitrogen .The other elements present are Carbon, hydrogen and Oxygen. On combustion of 0.085g of A in excess Oxygen,0.224g of carbon(IV)oxide and 0.0372g of water was formed. Determine the empirical formula of A (N=14.0, O=16.0 , C=12.0 , H=1.0)**

Mass of **N** in A = 5.2% x 0.085 **= 0.00442 g**

Mass of C in A = 12 x 0.224 = **0.0611g**

44

Mass of H in A = 2 x 0.0372 = **0.0041g**

18

Mass of O in A = 0.085g – 0.004442g = **0.0806g** (Mass of C,H,O)

=> 0.0611g + 0.0041g = **0.0652g (Mass of C,H)**

0.0806g (Mass of C,H,O)- 0.0652g (Mass of C,H) = **0.0154 g**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | Nitrogen | Carbon | Hydrogen | Oxygen |
| Symbol | N | C | H | O |
| Moles present = mass  Molar mass | 0.00442 g  14 | 0.0611g  12 | 0.0041g  1 | 0.0154 g 16 |
| Divide by the smallest value | 0.00032  0.00032 | 0.00509  0.00032 | 0.0041g  0.00032 | 0.00096  0.00032 |
| Mole ratios | 1 | 16 | 13 | 3 |

**Empirical formula = C16H13NO3**

**(d)Molar gas volume**

The volume occupied by one mole of all gases at the same temperature and pressure is a constant.It is:

**(i)** 24dm3/24litres/24000cm3 at **r**oom **t**emperature(25oC/298K)and **p**ressure(**r.t.p**).

i.e. 1mole of all gases =24dm3/24litres/24000cm3 atr.t.p

Examples

**1mole of O2** = 32g =6.0 x1023 particles**= 24dm3/24litres/24000cm3 at r.t.p**

**1mole of H2**= 2g =6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p**

**1mole of CO2**= 44g = 6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p**

**1mole of NH3**= 17g =6.0 x1023 particles **= 24dm3/24litres/24000cm3 at r.t.p**

**1mole of CH4** = 16g =6.0 x1023 particles **=24dm3/24litres/24000cm3 at r.t.p**

(**ii**)22.4dm3/22.4litres/22400cm3 at **s**tandard **t**emperature(0oC/273K) and **p**ressure(**s.t.p**)

i.e. 1mole of all gases =22.4dm3/22.4litres/22400cm3 ats.t.p

Examples

**1mole of O2** = 32g =6.0 x1023 particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of H2**= 2g =6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of CO2**= 44g = 6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of NH3**= 17g =6.0 x1023 particles**= 22.4dm3/22.4litres/22400cm3 at s.t.p**

**1mole of CH4** = 16g =6.0 x1023 particles **= 22.4dm3/22.4litres/22400cm3 at s.t.p**

The volume occupied by one mole of a gas at r.t.p or s.t.p is commonly called the **molar gas volume.** Whether the molar gas volume is at r.t.p or s.t.p must always be **specified**.

From the above therefore a less or more volume can be determined as in the examples below.

**Practice examples**

**1. Calculate the number of particles present in:**

(Avogadros constant =6.0 x1023mole-1 **)**

**(i) 2.24dm3 of Oxygen.**

22.4dm3 ->6.0 x1023

2.24dm3  **->** 2.24 x 6.0 x1023

22.4

=6.0 x1022 **molecules** = 2 x 6.0 x1022. = 1.2 x1023 **atoms**

**(ii) 2.24dm3 of Carbon(IV)oxide.**

22.4dm3 ->6.0 x1023

2.24dm3  **->** 2.24 x 6.0 x1023

22.4

=6.0 x1022 **molecules** = (CO2) = 3 x 6.0 x1022. = 1.8 x1023 **atoms**

**2. 0.135 g of a gaseous hydrocarbon X on complete combustion produces 0.41g of carbon(IV)oxide and 0.209g of water.0.29g of X occupy 120cm3 at room temperature and 1 atmosphere pressure .Name X and draw its molecular structure.(C=12.0,O= 16.O,H=1.0,1 mole of gas occupies 24dm3 at r.t.p)**

Molar mass CO2= 44 gmole-1√ Molar mass H2O = 18 gmole-1√

Molar mass X = 0.29 x (24 x 1000)cm3 = 58 gmole-1√

120cm3

Since a hydrocarbon is a compound containing Carbon and Hydrogen only. Then:

Mass of carbon in CO2 = Mass of C in CO2 x mass of CO2 =>

Molar mass of CO2

12 x 0.41 = **0.1118g**√

44

Mass of Hydrogen in H2O = Mass of C in H2Ox mass of H2O=>

Molar mass of H2O

2 x 0.209 = **0.0232g**√

18

|  |  |  |
| --- | --- | --- |
| Element | Carbon | Hydrogen |
| Symbol | C | H |
| Moles present = % composition  Molar mass | 0.g118  12 | 0.0232g√  1 |
| Divide by the smallest value | 0.0093  0.0093 | 0.0232  0.0093√ |
| Mole ratios | 1 x2 | 2.5x2 |
|  | 2 | 5√ |

Empirical formula is **C2H5**√

The molecular formular is thus determined :

**n**  = Relative formular mass = 58 = **2**√

Relative empirical formula 29

The molecular formula is (**C2H5** ) x 2 = **C4H10**.√

Molecule name **Butane**

Molecula structure

H H H H

H C C C C H√

H H H H

(e)**Gravimetric analysis**

Gravimetric analysis is the relationship between reacting masses and the volumes and /or masses of products. All reactants are in mole **ratios** to their products in accordance to their stoichiometric equation. Using the mole ration of reactants and products any volume and/or mass can be determined as in the examples:

1. Calculate the volume of carbon(IV)oxide at r.t.p produced when 5.0 g of calcium carbonate is strongly heated.(Ca=40.0, C= 12.0,O = 16.0,1 mole of gas =22.4 at r.t.p)

Chemical equation

CaCO3(s) -> CaO(s) + CO2(g)

Mole ratios 1: 1: 1

Molar Mass CaCO3 =100g

**Method 1**

100g CaCO3(s) -> 24dm3 CO2(g) at r.t.p

5.0 g CaCO3(s) -> 5.0 g x 24dm3 = **1.2dm3/1200cm3**

100g

**Method 2**

Moles of 5.0 g CaCO3(s) = 5.0 g = **0.05** moles

100 g

Mole ratio 1:1

Moles of CO2(g) = **0.05**moles

Volume of CO2(g) = 0.05 x 24000cm3 =**1200cm3 /1.2dm3**

**2. 1.0g of an alloy of aluminium and copper were reacted with excess hydrochloric acid. If 840cm3 of hydrogen at s.t.p was produced, calculate the % of copper in the alloy.(Al =27.0,one mole of a gas at s.t.p =22.4dm3 )**

Chemical equation

Copper does not react with hydrochloric acid

2Al(s) + 6HCl(aq) -> 2AlCl3(aq) + 3H2(g)

Method 1

3H2(g) = 3 moles x (22.4 x 1000)cm3 => 2 x 27 g Al

840cm3 => 840cm3 x 2 x 27 = **0.675**g of Aluminium

3 x 22.4 x 1000

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g =**0.325**g of copper

% copper = mass of copper x100% = **32.5%**

Mass of alloy

Method 2

Mole ratio 2Al: 3H2 = 2:3

Moles of Hydrogen gas = volume of gas => 840cm3 = **0.0375moles**

Molar gas volume 22400cm3

Moles of Al = 2/3 moles of H2 => 2/3x 0.0375moles = **0.025moles**

Mass of Al = moles x molar mass =>0.025moles x 27 = **0.675g**

Total mass of alloy – mass of aluminium = mass of copper

=> 1.0g - 0.675g = **0.325** g of copper

% copper = mass of copper x100% = **32.5%**

Mass of alloy

**(f)Gay Lussac’s law**

Gay Lussacs law states that **“when gases combine/react they do so in simple volume ratios to each other and to their gaseous products at constant/same temperature and pressure”**

Gay Lussacs law thus only apply to gases

Given the volume of one gas reactant, the other gaseous reactants can be deduced thus:

**Examples**

**1. Calculate the volume of Oxygen required to completely react with 50cm3 of Hydrogen.**

Chemical equation: 2H2 (g) + O2 (g) -> 2H2O(l)

Volume ratios 2 : 1 : 0

Reacting volumes 50cm3 : **25cm3**

**50cm3 of Oxygen is used**

**2. Calculate the volume of air required to completely reacts with 50cm3 of Hydrogen.(assume Oxygen is 21% by volume of air)**

Chemical equation: 2H2 (g) + O2 (g) -> 2H2O(l)

Volume ratios 2 : 1 : 0

Reacting volumes 50cm3 : **25cm3**

**50cm3 of Oxygen is used**

21% = 25cm3

100% = 100 x 25 =

21

3.If 5cm3 of a hydrocarbon CxHy burn in 15cm3 of Oxygen to form 10cm3 of Carbon(IV)oxide and 10cm3 of water vapour/steam, obtain the equation for the reaction and hence find the value of x and y in CxHy.

Chemical equation: CxHy (g) + O2 (g) -> H2O(g) + CO2(g)

Volumes 5cm3 : 15cm3 : 10cm3 : 10cm3

Volume ratios 5cm3 : 15cm3 : 10cm3 : 10cm3 (divide by lowest volume) **5 5 5 5**

**Reacting volume ratios** 1volume 3 volume 2 volume 2 volume

Balanced chemical equation: **CxHy (g) + 3O2 (g) -> 2H2O(g) + 2CO2(g)**

If “4H” are in 2H2O(g) the y=4

If “2C” are in 2CO2 (g) the x=2

Thus(i) chemical formula of hydrocarbon = **C2H4**

(ii) chemical name of hydrocarbon = **Ethene**

**4.100cm3 of nitrogen (II)oxide NO combine with 50cm3 of Oxygen to form 100cm3 of a single gaseous compound of nitrogen. All volumes measured at the same temperature and pressure. Obtain the equation for the reaction and name the gaseous product.**

Chemical equation: NO (g) + O2 (g) -> NOx

Volumes 100cm3 : 50cm3 : 100

Volume ratios 100cm3 : 50cm3 : 100cm3  (divide by lowest volume) **50 50 50**

**Reacting volume ratios** 2volume 1 volume 2 volume

Balanced chemical equation: 2 **NO (g) + O2 (g) -> 2NO x(g)**

Thus(i) chemical formula of the nitrogen compound = **2 NO2**

(ii) chemical name of compound = **Nitrogen(IV)oxide**

5.When 15cm3 of a gaseous hydrocarbon was burnt in 100cm3 of Oxygen ,the resulting gaseous mixture occupied70cm3 at room temperature and pressure. When the gaseous mixture was passed through, potassium hydroxide its volume decreased to 25cm3.

(a)What volume of Oxygen was used during the reaction.(1mk)

**Volume of Oxygen used =100-25 =75cm3√**

**(P was completely burnt)**

(b)Determine the molecular formula of the hydrocarbon(2mk)

**CxHy + O2 -> xCO2 + yH2O**

**15cm3 : 75cm3**

**15 15**

**1 : 3√**

**=> 1 atom of C react with 6 (3x2)atoms of Oxygen**

**Thus x = 1 and y = 2 => P has molecula formula CH4√**

**(g) Ionic equations**

An ionic equation is a chemical statement showing the movement of ions (cations and anions ) from reactants to products.

**Solids**, **gases** and **liquids** do not ionize/dissociate into free ions. **Only** ionic compounds in **aqueous/solution** or **molten** state ionize/dissociate into free cations and anions (**ions**)

An ionic equation is usually derived from a stoichiometric equation by using the following guidelines

Guidelines for writing ionic equations

1.Write the balanced stoichiometric equation

2.Indicate the state symbols of the reactants and products

3.**Split** into cations and anions all the reactants and products that exist in **aqueous** state.

4.**Cancel out** any cation and anion that appear on **both** the product and reactant side.

5. Rewrite the chemical equation. It is an ionic equation.

**Practice**

(**a**)Precipitation of an insoluble salt

All insoluble salts are prepared in the laboratory from double decomposition /precipitation. This involves mixing **two soluble** salts to form **one soluble** and **one insoluble** salt

1. When silver nitrate(V) solution is added to sodium chloride solution,sodium nitrate(V) solution and a white precipitate of silver chloride are formed.

Balanced stoichiometric equation

AgNO3(aq) + NaCl(aq) -> AgCl (**s**) + NaNO3 (aq)

Split reactants product existing in aqueous state as cation/anion

Ag+(aq) + NO3- (aq) + Na+(aq) + Cl-(aq) -> AgCl(**s**) + Na+(aq)+ NO3- (aq)

Cancel out ions appearing on reactant and product side

Ag+(aq) + NO3- (aq) + Na+(aq) + Cl-(aq) -> AgCl(**s**) + Na+(aq)+ NO3- (aq)

Rewrite the equation

**Ag+(aq) + Cl-(aq) -> AgCl(s) (**ionic equation**)**

2. When barium nitrate(V) solution is added to copper(II)sulphate(VI) solution, copper(II) nitrate(V) solution and a white precipitate of barium sulphate(VI) are formed.

Balanced stoichiometric equation

Ba(NO3)2(aq) + CuSO4(aq) -> BaSO4 (**s**) + Cu(NO3) 2 (aq)

Split reactants product existing in aqueous state as cation/anion

Ba2+(aq) + **2**NO3- (aq) + Cu2+(aq) + SO42-(aq) -> BaSO4 (**s**) + **2**NO3- (aq)+ Cu2+(aq)

Cancel out ions appearing on reactant and product side

Ba2+(aq) + 2NO3- (aq) +Cu2+ (aq) + SO42-(aq)-> BaSO4(**s**) + **2**NO3- (aq) + Cu2+(aq)

Rewrite the equation

**Ba2+(aq) + SO42-(aq) -> BaSO4(s) (**ionic equation**)**

**3.A yellow precipitate of Potassium Iodide is formed from the reaction of Lead(II)nitrate(v) and potassium iodide.**

Balanced stoichiometric equation

Pb(NO3)2(aq) + 2KI(aq) -> PbI2 (**s**) + 2KNO3  (aq)

Split reactants product existing in aqueous state as cation/anion

Pb2+(aq) + **2**NO3- (aq) + 2K +(aq) + 2I - (aq) -> PbI2 (**s**) + **2**NO3- (aq)+ 2K +(aq)

Cancel out ions appearing on reactant and product side

Pb2+(aq) + **2**NO3- (aq) + 2K +(aq) + 2I - (aq) -> PbI2 (**s**) + **2**NO3- (aq)+ 2K +(aq)

Rewrite the equation

**Pb2+(aq) + 2I- (aq) -> PbI2 (s) (**ionic equation**)**

(**b**)Neutralization

Neutralization is the reaction of an acid with a **soluble** base/alkali or **insoluble** base.

(i)Reaction of alkalis with acids

1.Reaction of nitric(V)acid with potassium hydroxide

Balanced stoichiometric equation

HNO3(aq) + KOH(aq) -> H2O (**l**) + KNO3  (aq)

Split reactants product existing in aqueous state as cation/anion

H+(aq) + NO3- (aq) + K +(aq) + OH - (aq) -> H2O (**l**) + NO3- (aq)+ K +(aq)

Cancel out ions appearing on reactant and product side

H+(aq) + NO3- (aq) + K +(aq) + OH - (aq) -> H2O (**l**) + NO3- (aq)+ K +(aq)

Rewrite the equation

**H+ (aq) + OH - (aq) -> H2O (l) (ionic equation)**

2.Reaction of sulphuric(VI)acid with ammonia solution

Balanced stoichiometric equation

H2SO4(aq) + 2NH4OH(aq) -> H2O (**l**) + (NH4) 2SO4  (aq)

Split reactants product existing in aqueous state as cation/anion

2H+(aq) + SO42- (aq) + 2NH4 +(aq)+ 2OH - (aq) ->2H2O (**l**) +SO42- (aq)+ 2NH4 + (aq)

Cancel out ions appearing on reactant and product side

2H+(aq) + SO42- (aq) + 2NH4 +(aq)+ 2OH - (aq) ->2H2O (**l**) +SO42- (aq)+ 2NH4 + (aq)

Rewrite the equation

2**H+ (aq) + 2OH - (aq) -> 2H2O (l)**

**H+ (aq) + OH - (aq) -> H2O (l) (ionic equation)**

3.Reaction of hydrochloric acid with Zinc hydroxide

Balanced stoichiometric equation

**2**HCl(aq) + Zn(OH)2 (**s**) -> 2H2O (**l**) + ZnCl 2  (aq)

Split reactants product existing in aqueous state as cation/anion

2H+(aq) + 2Cl- (aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + 2Cl- (aq)+ Zn 2+ (aq)

Cancel out ions appearing on reactant and product side

2H+(aq) + 2Cl- (aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + 2Cl- (aq)+ Zn 2+ (aq)

Rewrite the equation

2H+(aq) + Zn(OH)2 (**s**) ->2H2O (**l**) + Zn 2+ (aq) **(ionic equation)**

**(h)Molar solutions**

A molar solution is one whose concentration is known. The SI unit of concentration is **Molarity** denoted **M**.

Molarity may be defined as the number of moles of solute present in **one** cubic **decimeter** of solution.

One cubic decimeter is equal to **one litre** and also equal to **1000cm3**.

The higher the molarity the higher the concentration and the higher/more solute has been dissolved in the solvent to make one cubic decimeter/ litre/1000cm3 **solution.**

**Examples**

2M sodium hydroxide means 2 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

0.02M sodium hydroxide means 0.02 moles of sodium hydroxide solute is dissolved in enough water to make one cubic decimeter/ litre/1000cm3 uniform **solution** mixture of sodium hydroxide and water.

“2M” is **more concentrated** than“0.02M”.

**Preparation of molar solution**

Procedure

Weigh accurately 4.0 g of sodium hydroxide pellets into a 250cm3 volumetric flask.

Using a wash bottle add about 200cm3 of distilled water.

Stopper the flask.

Shake vigorously for three minutes.

Remove the stopper for a second then continue to shake for about another two minutes until **all** the solid has dissolved.

Add more water slowly upto **exactly** the 250 cm3 **mark**.

Sample questions

**1.Calculate the number of moles of sodium hydroxide pellets present in:**

**(i) 4.0 g.**

Molar mass of NaOH = (23 + 16 + 1) = 40g

Moles = Mass => 4.0 = **0.1**  /  **1.0 x 10 -1** moles

Molar mass 40

**(ii) 250 cm3 solution in the volumetric flask.**

Moles in 250 cm3 = **0.1 / 1.0 x 10 -1** moles

**(iii) one decimeter of solution**

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 x 10 -1 moles x 1000cm3 =

250cm3

= **0.4 M / 0.4 molesdm-3**

Method 2

250cm3 solution contain 1.0 x 10 -1 moles

1000cm3 solution = Molarity contain 1000 x 1.0 x 10 -1 moles

250 cm3

= **0.4 M / 0.4 molesdm-3**

**Theoretical sample practice**

**1. Calculate the molarity of a solution containing:**

**(i) 4.0 g sodium hydroxide dissolved in 500cm3 solution**

Molar mass of NaOH = (23 + 16 + 1) = 40g

Moles = Mass => 4.0 = **0.1**  /  **1.0 x 10 -1** moles

Molar mass 40

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 x 10 -1 moles x 1000cm3

500cm3

= **0.2 M / 0.2 molesdm-3**

Method 2

500 cm3 solution contain 1.0 x 10 -1 moles

1000cm3 solution = Molarity contain 1000 x 1.0 x 10 -1 moles

500 cm3

= **0.2 M / 0.2 molesdm-3**

**(ii) 5.3 g anhydrous sodium carbonate dissolved in 50cm3 solution**

Molar mass of Na2CO3 = (23 x 2 + 12 + 16 x 3) = 106 g

Moles = Mass => 5.3 = **0.05 / 5. 0 x 10-2**  moles

Molar mass 106

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.0 moles x 1000cm3 =

50cm3

=**1.0 M**

Method 2

50 cm3 solution contain 5.0 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10 -2 moles

50 cm3

= **1.0M / 1.0 molesdm-3**

**(iii) 5.3 g hydrated sodium carbonate decahydrate dissolved in 50cm3 solution**

Molar mass of Na2CO3.10H2O = (23 x 2 + 12 + 16 x 3 + 20 x 1 + 10 x 16) =286g

Moles = Mass => 5.3 = **0.0185 / 1.85 x 10 -2**  moles

Molar mass 286

Method 1

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.85 x 10 -2 moles x 1000cm3 =

50cm3

= **0.37 M/0.37 molesdm-3**

Method 2

50 cm3 solution contain 1.85 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 1.85 x 10 -2 moles

50 cm3

= **3.7 x 10**-1 **M / 3.7 x 10**-1 **molesdm-3**

**(iv) 7.1 g of anhydrous sodium sulphate(VI)was dissolved in 20.0 cm3 solution. Calculate the molarity of the solution.**

Method 1

20.0cm3 solution ->7.1 g

1000cm3 solution -> 1000 x 71 = 3550 g **dm-3**

20

Molar mass Na2SO4 = 142 g

Moles **dm-3** = Molarity = Mass 3550 = 2.5 M/ moles**dm-3**

Molar mass 142

Method 2

Molar mass Na2SO4 = 142 g

Moles = Mass => 7.1 = **0.05 / 5.0 x 10 -2**  moles

Molar mass 142

Method 2(a)

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 5.0 x 10 -2 moles x 1000cm3

20cm3

= **2.5 M/2.5 molesdm-3**

Method 2(b)

20 cm3 solution contain 5.0 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 5.0 x 10 -2 moles

20 cm3

= **2.5 M/2.5 molesdm-3**

**(iv) The density of sulphuric(VI) is 1.84gcm-3 Calculate the molarity of the acid.**

Method 1

1.0cm3 solution ->1.84 g

1000cm3 solution -> 1000 x 1.84 = 1840 g **dm-3**

1

Molar mass H2SO4 = 98 g

Moles **dm-3** = Molarity = Mass = 1840

Molar mass 98

= **18.7755 M**/ moles**dm-3**

Method 2

Molar mass H2SO4 = 98 g

Moles = Mass => 1.84 = **0.0188 / 1.88 x 10 -2**  moles

Molar mass 98

Method 2(**a**)

Moles in decimeters = **Molarity** = Moles x 1000cm3/1dm3

Volume of solution

=> 1.88 x 10 -2 moles x 1000cm3

1.0cm3

= **18.8M/18.8 molesdm-3**

Method 2(**b**)

20 cm3 solution contain 1.88 x 10 -2 moles

1000cm3 solution = Molarity contain 1000 x 1.88 x 10 -2 moles

1.0 cm3

= **18.8M/18.8 molesdm-3**

**2. Calculate the mass of :**

**(i) 25 cm3 of 0.2M sodium hydroxide solution(Na =23.0.O =16.0, H=1.0)**

Molar mass NaOH = **40g**

Moles in 25 cm3 = Molarity x volume => 0.2 x 25 = **0.005/5.0 x 10-3**moles

1000 1000

Mass of NaOH =Moles x molar mass = 5.0 x 10-3 x 40 = **0.2 g**

**(ii) 20 cm3 of 0.625 M sulphuric(VI)acid (S =32.0.O =16.0, H=1.0)**

Molar mass H2SO4 = **98g**

Moles in 20 cm3 = Molarity x volume=> 0.625 x 20 = **0.0125/1.25.0 x 10-3**moles

1000 1000

Mass of H2SO4 =Moles x molar mass => 5.0 x 10-3 x 40 = **0.2 g**

**(iii) 1.0 cm3 of 2.5 M Nitric(V)acid (N =14.0.O =16.0, H=1.0)**

Molar mass HNO3 = **63 g**

Moles in 1 cm3 = Molarity x volume => 2.5 x 1 = **0.0025 / 2.5. x 10-3**moles

1000 1000

Mass of HNO3 =Moles x molar mass => 2.5 x 10-3 x 40 = **0.1 g**

**3. Calculate the volume required to dissolve :**

**(a)(i) 0.25moles of sodium hydroxide solution to form a 0.8M solution**

Volume (in cm3) = moles x 1000 => 0.25 x 1000 = **312.5cm3**

Molarity 0.8

**(ii) 100cm3 was added to the sodium hydroxide solution above. Calculate the concentration of the solution.**

C1 x V1 = C2 x V2 where:

C1 = molarity/concentration before diluting/adding water

C2 = molarity/concentration after diluting/adding water

V1 = volume before diluting/adding water

V2 = volume after diluting/adding water

=> 0.8M x 312.5cm3 = C2 x (312.5 + 100**)**

C2 = 0.8M x 312.5cm3 = **0.6061M**

412.5

**(b)(ii) 0.01M solution containing 0.01moles of sodium hydroxide solution .**

Volume (in cm3) = moles x 1000 => 0.01 x 1000 = **1000 cm3**

Molarity 0.01

**(ii) Determine the quantity of water which must be added to the sodium hydroxide solution above to form a 0.008M solution.**

C1 x V1 = C2 x V2 where:

C1 = molarity/concentration before diluting/adding water

C2 = molarity/concentration after diluting/adding water

V1 = volume before diluting/adding water

V2 = volume after diluting/adding water

=> 0.01M x 1000 cm3 = 0.008 x V2

V2 = 0.01M x 1000cm3 = **1250cm3**

0.008

Volume added = 1250 - 1000 = **250cm3**

**(c)Volumetric analysis/Titration**

Volumetric analysis/Titration is the process of determining unknown concentration of one reactant from a known concentration and volume of another.

Reactions take place in simple mole ratio of reactants and products.

Knowing the concentration/ volume of one reactant, the other can be determined from the relationship:

**M1V1 = M2V2** where**:**

**n1 n2**

M1 = Molarity of 1st reactant

M2 = Molarity of 2nd reactant

V1 = Volume of 1st reactant

V1 = Volume of 2nd reactant

n1 = number of moles of 1st reactant from stoichiometric equation

n2 = number of moles of 2nd reactant from stoichiometric equation

Examples

**1.Calculate the molarity of MCO3 if 5.0cm3 of MCO3 react with 25.0cm3 of 0.5M hydrochloric acid.(C=12.0 ,O =16.0)**

Stoichiometric equation:MCO3(s) + 2HCl(aq) -> MCl2(aq) + CO2(g) + H2O(l)

Method 1

M1V1 = M2V2 -> M1 x 5.0cm3 = 0.5M x 25.0cm3

n1 n2 1 2

=> M1 = 0.5 x 25.0 x1 = **1.25M / 1.25 moledm-3**

5.0 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.5 x 25.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio MCO3 : HCl = 1:2

Moles MCO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity MCO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**2. 2.0cm3 of 0.5M hydrochloric acid react with 0.1M of M2CO3. Calculate the volume of 0.1M M2CO3 used.**

Stoichiometric equation:M2CO3 (aq) + 2HCl(aq) -> 2MCl (aq) + CO2(g) + H2O(l)

Method 1

M1V1 = M2V2 -> 0.5 x 2.0cm3 = 0.1M x V2 cm3

n1 n2 2 1

=> V2 = 0.5 x 2.0 x1 = **1.25M / 1.25 moledm-3**

0.1 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.5 x 2.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio M2CO3 : HCl = 1:2

Moles M2CO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity M2CO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**3. 5.0cm3 of 0.1M sodium iodide react with 0.1M of Lead(II)nitrate(V). Calculate(i) the volume of Lead(II)nitrate(V) used.**

**(ii)the mass of Lead(II)Iodide formed**

(Pb=207.0, I =127.0)

Stoichiometric equation: 2NaI(aq) + Pb(NO3)2(aq) -> 2NaNO3(aq) + PbI2(**s**)

(i)Volume of Lead(II)nitrate(V) used

Method 1

M1V1 = M2V2 -> 5 x 0.1cm3 = 0.1M x V2 cm3

n1 n2 2 1

=> V2 = 0.1 x 5.0 x 1 = **1.25M / 1.25 moledm-3**

0.1 x 2

Method 2

Moles of HCl used = molarity x volume

1000

=> 0.1 x 5.0 = **0.0125 /1.25 x 10-2moles**

1000

Mole ratio M2CO3 : HCl = 1:2

Moles M2CO3 = 0.0125 /1.25 x 10-2moles = **0.00625 / 6.25 x 10-3 moles**

2

Molarity M2CO3 = moles x 1000 => 0.00625 / 6.25 x 10-3  x 1000

Volume 5

= **1.25M / 1.25 moledm-3**

**4. 0.388g of a monobasic organic acid B required 46.5 cm3 of 0.095M sodium hydroxide for complete neutralization. Name and draw the structural formula of B**

Moles of NaOH used = molarity x volume

1000

=> 0.095 x 46.5 = **0.0044175 /4.4175 x 10-3moles**

1000

Mole ratio B: NaOH = 1:1

Moles B= **0.0044175 /4.4175 x 10-3moles**

Molar mass B = mass => 0.388

moles 0.0044175 /4.4175 x 10-3moles

= **87.8324 gmole-1**

X-COOH = 87.8324 where X is an alkyl group

X =87.8324- 42 = 42.8324=43

By elimination: CH3 = 15 CH3CH2 = 29 **CH3CH2 CH2  = 43**

**Molecula formula : CH3CH2 CH2COOH**

**Molecule name : Butan-1-oic acid**

**Molecular structure**

**H H H O**

**H C C C C O H**

H H H H

**5. 10.5 g of an impure sample containing ammonium sulphate (VI) fertilizer was warmed with 250cm3 of o.8M sodium hydroxide solution.The excess of the alkali was neutralized by 85cm3 of 0.5M hydrochloric acid. Calculate the % of impurities in the ammonium sulphate (VI)fertilizer. (N=14.0,S=32.0,O=16.0, H=1.0)**

Equation for neutralization

NaOH(aq) + HCl(aq) -> NaOH(aq) + H2O(l)

Mole ratio NaOH(aq):HCl(aq)= 1:1

Moles of HCl = Molarity x volume => 0.5 x 85 = **0.0425 moles**

1000 1000

Excess moles of NaOH(aq)= **0.0425 moles**

Equation for reaction with ammonium salt

2NaOH(aq) + (NH4) 2SO4(aq) -> Na 2SO4(aq) + 2NH3 (g)+ 2H2O(l)

Mole ratio NaOH(aq): (NH4) 2SO4(aq)= 2:1

Total moles of NaOH = Molarity x volume => 0.8 x 250 = **0.2 moles**

1000 1000

Moles of NaOH that reacted with(NH4) 2SO4 = 0.2 - 0.0425 = **0.1575moles**

Moles (NH4) 2SO4 = ½ x 0.1575moles = **0. 07875moles**

Molar mass (NH4) 2SO4= **132 gmole-1**

Mass of in impure sample = moles x molar mass =>0. 07875 x 132 = **10.395 g**

Mass of impurities = 10.5 -10.395 = **0.105 g**

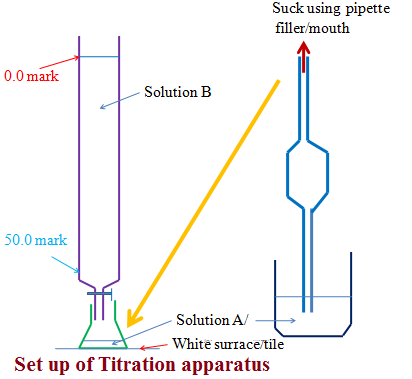
% impurities = 0.105 x 100 =  **1.0 %**

10.5

Practically volumetric analysis involves **titration**.

Titration generally involves filling a burette with known/unknown concentration of a solution then adding the solution to unknown/known concentration of another solution in a conical flask until there is complete reaction. If the solutions used are both colourless, an **indicator** is added to the conical flask. When the reaction is over, a **slight**/**little** excess of burette contents **change** the colour of the indicator. This is called the **end point**.

Set up of titration apparatus



The titration process involve involves determination of **titre.** The titre is the volume of burette contents/reading **before** and **after** the end point. Burette contents/reading **before** titration is usually called the **Initial** burette reading. Burette contents/reading **after** titration is usually called the **Final** burette reading. The titre value is thus a sum of the **Final** less **Initial** burette readings.

To reduce errors, titration process should be repeated at least once more.

The results of titration are recorded in a **titration table** as below

Sample titration table

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | 20.0 | 20.0 | 20.0 |
| Initial burette reading (cm3) | 0.0 | 0.0 | 0.0 |
| Volume of solution used(cm3) | 20.0 | 20.0 | 20.0 |

As **evidence** of a titration **actually** done examining body requires the candidate to record their burette readings before and after the titration.

For KCSE candidates burette readings **must** be recorded **in** a titration table in the **format** **provided** by the Kenya National Examination Council.

As **evidence** of all titration **actually** done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table **in the format** **provided**.

Calculate the average volume of solution used

24.0 + 24.0 + 24.0 = **24.0 cm3**

3

As evidence of understanding the degree of accuracy of burettes , all readings must be recorded to **a** decimal point.

As evidence of accuracy in carrying the out the titration , candidates value should be **within 0.2** of the **school** **value** .

The school value is the **teachers** readings presented to the examining body/council based on the concentrations of the solutions s/he presented to her/his candidates.

Bonus mark is awarded for averaged reading **within 0.1** school value as Final answer.

Calculations involved after the titration require candidates **thorough** practical and theoretical **practice mastery** on the:

(**i**)relationship among the mole, molar mass, mole ratios, concentration, molarity.

(**ii**) mathematical application of 1st principles.

Very useful information which candidates forget appears usually in the beginning of the question paper as:

“**You are provided with…**”

All calculation must be to the **4th decimal point** unless they divide fully to a lesser decimal point.

Candidates are expected to use a non programmable scientific calculator.

(**a**)Sample Titration Practice 1 **(Simple Titration)**

**You are provided with:**

0.1M sodium hydroxide solution A

Hydrochloric acid solution B

You are required to determine the concentration of solution B in moles per litre.

Procedure

Fill the burette with solution B. Pipette 25.0cm3 of solution A into a conical flask. Titrate solution A with solution B using phenolphthalein indicator to complete the titration table 1

Sample results Titration table 1

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution B used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

**1. Calculate the average volume of solution B used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

**2. How many moles of:**

**(i)solution A were present in 25cm3 solution.**

Moles of solution A = Molarity x volume = 0.1 x 25 = **2.5 x 10-3** moles

1000 1000

**(ii)solution B were present in the average volume.**

Chemical equation: NaOH(aq) + HCl(aq) -> NaCl(aq) + H2O(l)

Mole ratio 1:1 => Moles of A = Moles of B = **2.5 x 10-3** moles

**(iii) solution B in moles per litre.**

Moles of B per litre = moles x 1000 = 2.5 x 10-3 x 1000 = **0.1M**

Volume 20

(**b**)Sample Titration Practice 2 **(Redox Titration)**

**You are provided with:**

Acidified Potassium manganate(VII) solution A

0.1M of an iron (II)salt solution B

8.5g of ammonium iron(II)sulphate(VI) crystals(NH4)2 SO4FeSO4.xH2O solid C

You are required to

(i)standardize acidified potassium manganate(VII)

(ii)determine the value of x in the formula (NH4)2 SO4FeSO4.xH2O.

**Procedure 1**

Fill the burette with solution A. Pipette 25.0cm3 of solution B into a conical flask. Titrate solution A with solution B until a pink colour just appears.

Record your results to complete table 1.

Table 1:**Sample results**

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution A used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

**1. Calculate the average volume of solution A used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

**2. How many moles of:**

**(i)solution B were present in 25cm3 solution.**

Moles of solution A = Molarity x volume = 0.1 x 25 = **2.5 x 10-3** moles

1000 1000

**(ii)solution A were present in the average volume. Assume one mole of B react with five moles of B**

Mole ratio A : B = 1:5

=> Moles of A = Moles of B = 2.5 x 10-3 moles = **5.0 x 10 -4 moles**

5 5

**(iii) solution B in moles per litre.**

Moles of B per litre = moles x 1000 = 2.5 x 10-3 x 1000

Volume 20

= **0.025** M /moles per litre /moles l-1

**Procedure 2**

Place all the solid C into the 250cm3 volumetric flask carefully. Add about 200cm3 of distilled water. Shake to dissolve. Make up to the 250cm3 of solution by adding more distilled water. Label this solution C. Pipette 25cm3 of solution C into a conical flask, Titrate solution C with solution A until a permanent pink colour just appears. Complete table 2.

Table 2:**Sample results**

|  |  |  |  |
| --- | --- | --- | --- |
| Titration number | 1 | 2 | 3 |
| Final burette reading (cm3) | **20.0** | **20.0** | **20.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of solution A used(cm3) | **20.0** | **20.0** | **20.0** |

Sample worked questions

**1. Calculate the average volume of solution A used**

Average titre = Titre 1 + Titre 2 +Titre 3 => ( 20.0 +20.0 +20.0 ) = **20.0cm3**

3 3

**2. How many moles of:**

**(i)solution A were present inin the average titre.**

Moles of solution A = Molarity x volume = 0.025 x 20 = **5.0 x 10-4** moles

1000 1000

**(ii)solution C in 25cm3 solution given the equation for the reaction:**

**MnO4- (aq) + 8H+(aq) + 5Fe2+ (aq) -> Mn2+(aq) + 5Fe3+(aq) + 4H2O(l)**

Mole ratio MnO4- (aq): 5Fe2+ (aq) = 1:5 => Moles of 5Fe2+ (aq) = Moles of MnO4- (aq) = 5.0 x 10-4 moles = **1.0 x 10 -4 moles**

5 5

**(iii) solution B in 250cm3.**

Moles of B per litre = moles x 250 = 1.0 x 10 -4  x 250 = **1.0 x 10 -3 moles** Volume 25

3. Calculate the molar mass of solid C and hence the value of x in the chemical formula (NH4)2SO4FeSO4.xH2O.

(N=14.0, S=32.0, Fe=56.0, H=1.0 O=16.0)

Molar mass = mass perlitre = 8.5 =  **8500 g**

Moles per litre 1.0 x 10 -3 moles

NH4)2SO4FeSO4.xH2O = 8500

284 + 18x =8500

8500 - 284 = 8216 = 18x =  **454.4444**

18 18

x = **454** (whole number)

**(**c**)**Sample Titration Practice 3 **(Back titration)**

You are provided with:

(i)an impure calcium carbonate labeled M

(ii)Hydrochloric acid labeled solution N

(iii)solution L containing 20g per litre sodium hydroxide.

You are required to determine the concentration of N in moles per litre and the % of calcium carbonate in mixture M.

**Procedure 1**

Pipette 25.0cm3 of solution L into a conical flask. Add 2-3 drops of phenolphthalein indicator. Titrate with dilute hydrochloric acid solution N and record your results in table 1(4mark)

Sample Table 1

|  |  |  |
| --- | --- | --- |
| 1 | 2 | 3 |
| Final burette reading (cm3) | **6.5** | **6.5** | **6.5** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of N used (cm3) | **6.5** | **6.5** | **6.5** |

Sample questions

**(a) Calculate the average volume of solution N used**

6.5 + 6.5 + 6.5 = **6.5** cm3

3

**(b) How many moles of sodium hydroxide are contained in 25cm3of solution L**

Molar mass NaOH =**40**g

Molarity of L = mass per litre => 20 = **0.5**M

Molar mass NaOH 40

Moles NaOH in 25cm3 = molarity x volume => 0.5M x 25cm3 =  **0.0125** moles

1000 1000

**(c)Calculate:**

**(i)the number of moles of hydrochloric acidthat react with sodium hydroxide in (b)above.**

Mole ratio NaOH : HCl from stoichiometric equation= **1:1**

Moles HCl =Moles NaOH => **0.0125** moles

**(ii)the molarity of hydrochloric acid solution N.**

Molarity = moles x 1000 => 0.0125 moles x 1000 =**1.9231**M/moledm-3

6.5 6.5

**Procedure 2**

Place the 4.0 g of M provided into a conical flask and add 25.0cm3 of the dilute hydrochloric acid to it using a clean pipette. Swirl the contents of the flask vigorously until effervescence stop.Using a 100ml measuring cylinder add 175cm3 distilled waterto make up the solution up to 200cm3.Label this solution K.Using a clean pipettetransfer 25.0cm3 of the solution into a clean conical flask and titrate with solution L from the burette using 2-3 drops of methyl orange indicator.Record your observations in table 2.

Sample Table 2

|  |  |  |
| --- | --- | --- |
| 1 | 2 | 3 |
| Final burette reading (cm3) | **24.5** | **24.5** | **24.5** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of N used (cm3) | **24.5** | **24.5** | **24.5** |

Sample calculations

**(a)Calculate the average volume of solution L used(1mk)**

24.5 + 24.5 + 24.5 = **24.5**cm3

3

**(b)How many moles of sodium hydroxide are present in the average volume of solution L used?**

Moles = molarity x average burette volume => 0.5 x 24.5

1000 1000

= **0.01225 /1.225 x 10-2** moles

**(c) How many moles of hydrochloric acid are present in the original 200cm3 of solution K?**

Mole ratio NaOH: HCl = 1:1 => moles of HCl = **0.01225 /1.225 x 10-2** moles

Moles in 200cm3 = 200cm3 x 0.01225 /1.225 x 10-2moles

25cm3(volume pipetted)

=**0.49 /4.9 x 10-1**moles

**(d)How many moles of hydrochloric acid were contained in original 25 cm3 solution N used**

Original moles = Original molarity x pipetted volume =>

1000cm3

1.9231M/moledm-3 x 25 =  **0.04807/4.807 x 10-2** moles

1000

**(e)How many moles of hydrochloric acid were used to react with calcium carbonate present?**

Moles that reacted = original moles –moles in average titre =>

0.04807/4.807 x 10-2moles - 0.01225 /1.225 x 10-2moles

= **0.03582/3.582 x 10 -2**moles

**(f)Write the equation for the reaction between calcium carbonate and hydrochloric acid.**

CaCO3(s) + 2HCl(aq) -> CaCl2(aq) + CO2(g) + H2O(l)

**(g)Calculate the number of moles of calcium carbonate that reacted with hydrochloric acid.**

From the equation CaCO3(s):2HCl(aq) = 1:2

=> Moles CaCO3(s) = 1/2moles HCl

= 1/2 x 0.03582/3.582 x 10 -2 moles

= **0.01791 /1.791 x 10-2**moles

**(h)Calculate the mass of calcium carbonate in 4.0g of mixture M (Ca=40.0,O = 16.0,C=12.0)**

Molar mass CaCO3 = 100g

Mass CaCO3 = moles x molar mass => 0.01791 /1.791 x 10-2moles x 100g

= **1.791**g

**(i)Determine the % of calcium carbonate present in the mixture**

% CaCO3 = mass of pure x 100% => 1.791g x 100% = **44.775**%

Mass of impure 4.0

(d)Sample titration practice 4 (**Multiple titration**)

You are provided with:

(i)sodium L containing 5.0g per litre of a dibasic organic acid H2X.2H2O.

(ii)solution M which is acidified potassium manganate(VII)

(iii)solution N a mixture of sodium ethanedioate and ethanedioic acid

(iv)0.1M sodium hydroxide solution P

(v)1.0M sulphuric(VI)

You are required to:

(i)standardize solution M using solution L

(ii)use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

**Procedure 1**

Fill the burette with solution M. Pipette 25.0cm3 of solution L into a conical flask. Heat this solution to about 70oC(**but not to boil**).Titrate the hot solution L with solution M until a permanent pink colour just appears .Shake thoroughly during the titration. Repeat this procedure to complete table 1.

Sample Table 1

|  |  |  |
| --- | --- | --- |
| 1 | 2 | 3 |
| Final burette reading (cm3) | **24.0** | **24.0** | **24.0** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of N used (cm3) | **24.0** | **24.0** | **24.0** |

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

24.0 + 24.0 + 24.0 = **24.0**cm3

3

**(b)Given that the concentration of the dibasic acid is 0.05molesdm-3.determine the value of x in the formula H2X.2H2O (H=1.0,O=16.0)**

Molar mass H2X.2H2O= mass per litre => 5.0g/litre = **100**g

Moles/litre 0.05molesdm-3

H2X.2H2O =100

X = 100 – ((2 x1) + 2 x (2 x1) + (2 x 16) => 100 – 34 = **66**

**(c) Calculate the number of moles of the dibasic acid H2X.2H2O.**

Moles = molarity x pipette volume => 0.5 x 25 = **0.0125/1.25 x10 -2** moles

1000 1000

**(d)Given the mole ratio manganate(VII)(MnO4-): acid H2X is 2:5, calculate the number of moles of manganate(VII) (MnO4-) in the average titre.**

Moles H2X = 2/5 moles of MnO4-

=> 2/5 x 0.0125/1.25 x10 -2 moles

= **0.005/5.0 x 10 -3**moles

**(e)Calculate the concentration of the manganate(VII)(MnO4-) in moles per litre.**

Moles per litre/molarity = moles x 1000

average burette volume

=>0.005/5.0 x 10 -3moles x 1000 =  **0.2083** molesl-1/M

24.0

**Procedure 2**

With solution M still in the burette ,pipette 25.0cm3 of solution N into a conical flask. Heat the conical flask containing solution N to about 70oC.Titrate while hot with solution M.Repeat the experiment to complete table 2.

Sample Table 2

|  |  |  |
| --- | --- | --- |
| 1 | 2 | 3 |
| Final burette reading (cm3) | **12.5** | **12.5** | **12.5** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of N used (cm3) | **12.5** | **12.5** | **12.5** |

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

12.5 + 12.5 + 12.5 =**12.5**cm3

3

**(b)Calculations:**

**(i)How many moles of manganate(VII)ions are contained in the average volume of solution M used?**

Moles = molarity of solution M x average burette volume

1000

=> 0.2083 molesl-1/ M x 12.5 = **0.0026 / 2.5 x 10-3** moles

1000

**(ii)The reaction between manganate(VII)ions and ethanedioate ions that reacted with is as in the equation:**

**2MnO4- (aq) + 5C2O42- (aq) + 16H+ (aq) -> 2Mn2+(aq) + 10CO2(g) + 8H2O(l)**

**Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M.**

From the stoichiometric equation,mole ratio MnO4- (aq): C2O42- (aq) = 2:5

=> moles C2O42- = 5/2 moles MnO4- => 5/2 x 0.0026 / 2.5 x 10-3 moles

= **0.0065 /6.5 x10-3** moles

**(iii)Calculate the number of moles of ethanedioate ions contained in 250cm3 solution N.**

25cm3 pipette volume -> 0.0065 /6.5 x10-3 moles

250cm3 ->

0.0065 /6.5 x10-3 moles x 250 = **0.065 / 6.5 x10-2** moles

25

**Procedure 3**

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25cm3 of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to complete table 3.

Sample Table 2

|  |  |  |
| --- | --- | --- |
| 1 | 2 | 3 |
| Final burette reading (cm3) | **24.9** | **24.9** | **24.9** |
| Initial burette reading (cm3) | **0.0** | **0.0** | **0.0** |
| Volume of N used (cm3) | **24.9** | **24.9** | **24.9** |

Sample calculations

**(a)Calculate the average volume of solution L used (1mk)**

24.9 + 24.9 + 24.9 =  **24.9 cm3**

3

**(b)Calculations:**

**(i)How many moles of sodium hydroxide solution P were contained in the average volume?**

Moles = molarity of solution P x average burette volume

1000

=> 0.1 molesl-1 x 24.9 = **0.00249 / 2.49 x 10-3** moles

1000

**(ii)Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:**

**2NaOH(aq) + H2C2O4 (aq) -> Na2C2O4(g) + 2H2O(l)**

**Calculate the number of moles of ethanedioic acid that were used in the reaction**

From the stoichiometric equation,mole ratio NaOH(aq): H2C2O4 (aq) = 2:1

=> moles H2C2O4 = 1/2 moles NaOH => 1/2 x 0.00249 / 2.49 x 10-3 moles

= **0.001245/1.245 x10-3** moles.

**(iii)How many moles of ethanedioic acid were contained in 250cm3 of solution N?**

25cm3 pipette volume -> 0.001245/1.245 x10-3 moles

250cm3 ->

0.001245/1.245 x10-3 moles x 250 = **0.01245/1.245 x10-2** moles

25

**(iii)Determine the % by mass of sodium ethanedioate in the micture (H= 1.0,O=16.0,C=12.0 and total mass of mixture =2.0 g in 250cm3 solution)**

Molar mass H2C2O4 = **90.0**g

Mass of H2C2O4 in 250cm3 = moles in 250cm3 x molar mass H2C2O4

=>0.01245/1.245 x10-2 moles x 90.0

= **1.1205**g

% by mass of sodium ethanedioate

=(Mass of mixture - mass of H2C2O4) x 100%

Mass of mixture

=> 2.0 - 1.1205 g =  **43.975**%

2.0

**Note**

**(i) L is 0.05M Oxalic acid**

**(ii) M is 0.01M KMnO4**

**(iii) N is 0.03M oxalic acid(without sodium oxalate)**

**Practice example 5.(Determining equation for a reaction)**

You are provided with

-0.1M hydrochloric acid solution A

-0.5M sodium hydroxide solution B

You are to determine the equation for thereaction between solution A and B

**Procedure**

Fill the burette with solution A.Using a pipette and pipette filler transfer 25.0cm3 of solution B into a conical flask.Add 2-3 drops of phenolphthalein indicator.Run solution A into solution B until a permanent pink colour just appears.Record your results in Table 1.Repeat the experiment to obtain three concordant results to complete Table 1

Table 1(Sample results)

|  |  |  |  |
| --- | --- | --- | --- |
| Titration | 1 | 2 | 3 |
| Final volume(cm3) | **12.5** | **25.0** | **37.5** |
| Initial volume(cm3) | **0.0** | **12.5** | **25.0** |
| Volume of solution A used(cm3) | **12.5** | **12.5** | **12.5** |

**Sample questions**

**Calculate the average volume of solution A used.**

12.5+12.5+12.5 = **12.5cm3**

**3**

**Theoretical Practice examples**

**1. 1.0g of dibasic acid HOOC(CH2)xCOOH was dissolved in 250cm3 solution. 25.0 cm3 of this solution reacted with 30.0cm3 of 0.06M sodium hydroxide solution. Calculate the value of x in HOOC(CH2)xCOOH. (C=12.0,H=1.0,O=16.)**

Chemical equation

2NaOH(aq) + H**2**X(aq) -> Na**2**X (aq) + 2H**2**O(aq)

Mole ratio NaOH(aq) :H**2**X(aq) = 2:1

Method 1

Ma Va = na => Ma x 25.0 = 1 => Ma =0.06 x 30.0 x1

Mb Vb = nb 0.06 x 30.0 2 25.0 x 2

Molarity of acid = **0.036M/Mole l-1**

Mass of acid per lite = 1.0 x1000 = **4.0 g/l**

250

0.036M/ Mole l-1  -> 4.0 g /l

1 mole= molar mass of HOOC(CH2)xCOOH = 4.0 x 1 = **111.1111** g

0.036

Molar mass (CH2)x = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

(CH2)x = 14x = 21.1111 =  **1.5 = 1 (whole number)**

14

Method 2

Moles of sodium hydroxide = Molarity x volume = 0.06 x 30 = **1.8 x 10 -3moles**

1000

Moles of Hydrochloric acid = 1/2 x 1.8 x 10 -3moles = 9.0 x10 -4moles

Molarity of Hydrochloric acid = moles x 1000 = 9.0 x10 -4moles x1000

Volume 25

Molarity of acid = **0.036M/Mole l-1**

Mass of acid per lite = 1.0 x1000 = **4.0 g/l**

250

0.036M/ Mole l-1  -> 4.0 g /l

1 mole= molar mass of HOOC(CH2)xCOOH = 4.0 x 1 = **111.1111** g

0.036

Molar mass (CH2)x = 111.1111 – (HOOCCOOH = 90.0) **= 21.1111**

(CH2)x = 14x = 21.1111 = **1.5 = 1 (whole number)**

14

**2. 20.0cm3 of 0.05 M acidified potassium manganate(VII)solution oxidized 25.0cm3 of Fe2+(aq) ions in 40.0g/l of impure Iron (II)sulphate(VI) to Fe3+(aq) ions. Calculate the percentage impurities in the Iron (II)sulphate(VI).**

**MnO4- (aq) + 8H+(aq)+ 5Fe2+(aq)-> 5Fe3+(aq) + Mn2+(aq) + 4H2O(aq)**

**Fe=56.0,S= 32.0, O=16.0).**

Moles of MnO4- (aq) = Molarity x volume = 0.05 x 20.0 = **0.001 Moles**

1000 1000

Mole ratio MnO4- (aq): 5Fe2+(aq)= 1:5

Moles 5Fe2+(aq) **=** 5 x0.001  **= 0.005 Moles**

Moles of 5Fe2+(aq) per litre/molarity = Moles x 1000 = 0005 x 1000

Volume 25.0

**= 0.2 M/ Moles/litre**

Molar mass =FeSO4=**152 g**

Mass of in the mixture = Moles x molar mass => 0.2 x 152 = **30.4** g

Mass of impurity = 40.0 – 30.4 =**9.6 g**

% impurity **=** 9.6 g x100 = **24.0 % impurity**

40.0

**3.9.7 g of a mixture of Potassium hydroxide and Potassium chloride was dissolved to make one litre solution.20.0cm3 of this solution required 25.0cm3 of 0.12M hydrochloric acid for completed neutralization. Calculate the percentage by mass of Potassium chloride.(K=39.0,Cl= 35.5)**

Chemical equation

KOH(aq) + HCl(aq) -> KCl(aq) + H2O(l)

Moles of HCl = Molarity x volume => 0.12 x 25.0 = **0.003/3.0 x 10 -3** moles

1000 1000

Mole ratio KOH(aq) : HCl(aq) -= 1:1

Moles KOH =**0.003/3.0 x 10 -3** moles

Method 1

Molar mass KOH =**56.0**g

Mass KOH in 25cm3 =0.003/3.0 x 10 -3 moles x56.0 = 0.168g

Mass KOH in 1000cm3/1 litre = 0.168 x1000= **8.4 g/l**

20

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = 1.3 x 100 = **13.4021%**

9.7

Method 2

Moles KOH in 1000cm3 /1 litre = Moles in 20cm3 x 1000 =>0.003 x 1000

20 20

=**0.15M/Moles /litre**

Molar mass KOH =**56.0**g

Mass KOH in 1000/1 litre = 0.15M/Moles /litre x 56.0 = **8.4g/l**

Mass of KCl = 9.7g - 8.4g = **1.3** g

% of KCl = 1.3 x 100 = **13.4021%**

9.7

4.**A certain carbonate, GCO3, reacts with dilute hydrochloric acid according to the equation given below:**

**GCO3(s) + 2HCl(aq) -> GCl2 (aq) + CO2 (g) + H2O(l)**

**If 1 g of the carbonate reacts completely with 20 cm3 of 1 M hydrochloric acid ,calculate the relative atomic mass of G (C = 12.0 = 16.0)**

Moles of HCl = Molarity x volume=> 1 x20 = **0.02 moles**

1000 1000

Mole ratio HCl; GCO3  = 2:1

Moles of GCO3= 0.02 moles = **0.01moles**

2

Molar mass of GCO3  = mass => 1 = **100** g

moles 0.01moles

G= GCO3 - CO3 =>100g – (12+ 16 x3 = 60) = **40**(**no units**)

**5. 46.0g of a metal carbonate MCO3 was dissolved 160cm3 of 0.1M excess hydrochloric acid and the resultant solution diluted to one litre.25.0cm3 of this solution required 20.0cm3 of 0.1M sodium hydroxide solution for complete neutralization. Calculate the atomic mass of ‘M’**

Equation

Chemical equation

NaOH(aq) + HCl(aq) -> KCl(aq) + H2O(l)

Moles of NaOH = Molarity x volume=> 0.1 x20 = **0.002 moles**

1000 1000

Mole ratio HCl; NaOH = 1:1

Excess moles of HCl =  **0.002 moles**

25cm3 -> 0.002 moles

1000cm3 -> 1000 x 0.002 = **0.08moles**

25cm3

Original moles of HCl = Molarity x volume => 1M x 1litre = **1.0 moles**

Moles of HCl reacted with MCO3 = **1.0 - 0.08 moles** =  **0.92moles**

**Chemical equation**

MCO3(s) + 2HCl(aq) -> MCl2 (aq) + CO2 (g) + H2O(l)

Mole ratio MCO3(s) : HCl(aq) =1:2

Moles of MCO3  = 0.92moles => **0.46moles**

2

Molar mass of MCO3= mass => 46g = **100 g**

moles 0.46moles

M= MCO3 - CO3 =>100g – (12+ 16 x3 = 60) = **40**

**6. 25.0cm3 of a mixture of Fe2+ and Fe3+ ions in an aqueous salt was acidified with sulphuric(VI)acid then titrated against potassium manganate(VI).The salt required 15cm3 ofe0.02M potassium manganate(VI) for complete reaction.**

**A second 25cm3 portion of the Fe2+ and Fe3+ ion salt was reduced by Zinc then titrated against the same concentration of potassium manganate(VI).19.0cm3 of potassium manganate(VI)solution was used for complete reaction. Calculate the concentration of Fe2+ and Fe3+ ion in the solution on moles per litre.**

Mole ratio Fe2+ :Mn04- = 5:1

Moles Mn04- used = 0.02 x 15 = **3.0 x 10-4** moles

1000

Moles **Fe2+ =**  3.0 x 10-4 moles **= 6.0 x 10-5** moles

5

Molarity of **Fe2+ = 6.0 x 10-4** moles x 1000 =  **2.4 x 10-3 moles l-1**

25

Since Zinc reduces **Fe3+** to  **Fe2+**  in the mixture:

Moles Mn04- that reacted with all **Fe2+**= 0.02 x 19 = **3.8 x 10-4** moles

1000

Moles of all **Fe2+** = 3.8 x 10-4 moles = 7.6 x 10-5 moles

5

Moles of Fe3+  = 3.8 x 10-4 - 6.0 x 10-5 = **1.6 x 10-5** moles

Molarity of **Fe3+ =** 1.6 x 10-5 moles x 1000 = **4.0 x 10-4 moles l-1**

25