**CHEMISTRY SS 1**

**SECOND TERM**

**SCHEME OF WORK**

**WEEK TOPIC**

1. Revision/Introduction to the mole concept: Molar volume of gases, Avogadro’s number, Percentage of element in a compound.

Writing and Balancing Chemical Equations.

2-3. Stoichiometry of Reactions: Calculation of Masses of Reactants and Products,

Calculation of Volume of Reacting Gases.

Empirical and Molecular formulae.

4-5. Chemical Laws and their Verification: Law of Conservation of Mass, Law of Constant Composition, Law of Multiple Proportion.

6-7. Chemical Combinations: Electrovalent Bond: Properties of Electrovalent

Compounds, Covalent Bond: Properties of Covalent Compounds. Other Types of Bonding

8-9. TheKinetic theoryof Matter and the Gas Laws: Boyle’s Law, Charles’ Law, Ideal Gas Equation, Dalton’s Law of Partial Pressure.

10. Avogadro’s law, Gay-Lussac’s Law of Combining Volumes, Graham’s Law of Diffusion.

**REFERENCE BOOKS**

* New Chemistry for Senior Secondary School by Osei Yaw Ababio; U.T.M.E Past Questions and Answers.
* Practical Chemistry for Senior Secondary Schools by Godwin Ojokuku
* Outline Chemistry for Schools & Colleges by Ojiodu C.C.
* Chemistry Pass Questions for S.S.C.E and UTME.

**WEEK ONE**

**TOPIC:INTRODUCTION TO THE MOLE CONCEPT**

**CONTENT**

* Relative atomic mass
* Relative molecular mass
* Molar volume of gases
* Percentage of an element in a compound

**THE MOLE**

A mole is a number of particles of a substance which may be atoms, ions, molecules or electrons. This number of particles is approximately 6.02 x 1023 in magnitude and is known as Avogadro’s number of particles.

The mole is defined as the amount of a substance which contains as many elementary units as there are atoms in 12g of Carbon-12.

**RELATIVE ATOMIC MASS**

The relative atomic mass of an element is the number of time the average mass of one atom of that element is heavier than one twelfth the mass of one atom of Carbon-12. It indicates the mass of an atom of an element. For e.g, the relative atomic mass of hydrogen, oxygen, carbon, sodium and calcium are 1, 16, 12, 23, and 40 respectively.

The atomic mass of an element contains the same number of atoms which is 6.02 x 1023atoms; 1 mole of hydrogen having atomic mass of 2.0g contains 6.02 x 1023 atoms.

**EVALUATION**

1. Define relative atomic mass of an element
2. State the relative atomic mass of the following elements: potassium, chlorine, silver, lead, phosphorus and nitrogen

**RELATIVE MOLECULAR MASS**

The relative molecular mass of an element or compound is the number of times the average mass of one molecule of it is heavier than one-twelfth the mass of one atom of Carbon-12

It is the sum of the relative atomic masses of all atoms in one molecule of that substance. It is also called the formula mass. The formula mass refers not only to the relative mass of a molecule but also that of an ion or radical.

**CALCULATION**

Calculate the relative molecular mass of:

1. Magnesium chloride
2. Sodium hydroxide
3. Calcium trioxocarbonate

[Mg=24, Cl=35.5, Na=23, O=16, H=1, Ca=40,C=12]

**Solution:**

1. MgCl 2 = 24 + 35.5x2 = 24 + 71 = 95gmol-1
2. NaOH = 23 + 16 + 1 = 40gmol-1
3. CaCO3 = 40 + 12 +16x3 = 100gmol-1

**EVALUATION**

1. What is relative molecular mass of a compound?
2. Calculate the relative molecular mass of (a) NaNO3 (b) CuSO4.5H2O

**MOLAR VOLUME OF GASES**

The volume occupied by 1 mole of a gas at standard conditions of temperature and pressure (s.t.p) is 22.4 dm3. Thus 1 mole of oxygen gas of molar mass 32.0gmol-1 occupies a volume of 22.4dm3 at s.t.p and 1 mole of helium gas of molar mass 4.0gmol-1 occupies a volume of 22.4 dm3 at s.t.p.

Note: When the conditions of temperature and pressure are altered, the molar volume will also change. Also, standard temperature = 273K and standard pressure = 760mmHg.

**RELATIONSHIP BETWEEN QUANTITIES**

Molar mass = mass(g) i.e. M = m gmol-1

Amount (moles) n

Note: Amount = Number of moles

Molar volume of gas = volume ( cm3 or dm3) i.e. Vm = v dm3mol-1

Amount (mole)      n

Amount = Reacting mass (g)

Molar mass (gmol-1)

Also, Amount of substance = Number of particles

Avogadro’s constant

But, Avogadro’s constant = 6.02 x 1023

Combining the two expressions:

Reacting mass = Number of particles

Molar mass 6.02 x 1023

**CALCULATIONS**

1. What is the mass of 2.7 mole of aluminium (Al=27)?

**Solution:**

Amount = Reacting mass

Molar mass

Reacting mass = Amount x Molar mass

= 2.7mole x 27 gmol-1 = 72.9g.

2. What is the number of oxygen atoms in 32g of the gas (O=16, NA = 6.02 x 1023)?

**Solution:**

Reacting mass= Number of atoms

       Molar mass 6.02 x 1023

Number of atoms = Reacting mass x 6.02 x 1023

Molar mass

Molar mass of O2 = 16x2 =32gmol-1

Number of atoms = 32g x 6.02 x 1023

32gmol-1

= 6.02 x 1023

The number of oxygen atoms is 6.02 x 1023

**EVALUATION**

1. Define the molar volume of a gas

2. How many molecules are contained in 1.12dm3 of hydrogen gas at s.t.p?

**PERCENTAGE OF AN ELEMENT IN A COMPOUND**

The percentage composition of an atom in a compound is the amount of the atom expressed in percentage.

Percentage of an element in a compound = Mass of element in the compound x 100

Molar mass of compound 1

**CALCULATIONS**

1. What is the percentage by mass of nitrogen in NH4NO3 ( H=1, N=14, 0=16)?

**Solution:**

Molar mass of NH4NO3 = 14x2 + 1x4 + 16x3 = 80gmol-1

Percentage by mass of N2 = Mass of N2 x 100

Molar mass of NH4NO3 1

= 28 x 100= 35%

80 1

2. Calculate the percentage by mass of water of crystallization in MgSO4.7H2O

(Mg=24, S=32, 0=16, H=1)

**Solution:**

Molar mass of MgSO4.7H20 = 24 + 32 + 16x4 + 9(2+16) = 246gmol-1

7 moles of water of crystallization = 126g

Percentage by mass of water = Mass of H2O x 100

Molar mass of MgSO4.7H2O 1

= 126g x 100

246gmol-11

= 51.2%

**GENERAL EVALUATION**

1. What is the number of molecules in 6.4g of SO2 (NA = 6.02 X 1023)?
2. What is the volume in cm3 of 2.2g of CO2 at s.t.p ( C=12, O=16)?
3. Determine the percentage by mass of oxygen in Al2(SO4).2H2O.( Al=27, S=32, O=16, H=1)

**WEEKEND ASSIGNMENT**

1. What is the relative atomic mass of potassium A. 40 B. 39 C. 32 D. 24
2. An element with relative atomic mass 108 is A. Ca B. Cl C. Ag D. Al
3. Modern standard element with which chemist define relative atomic mass is A.12C B.C13 C.3H D.16O
4. Calculate the relative molecular mass of CH3COOH.A. 60gmol-1B. 70gmol-1C. 80gmol-1D. 90gmol-1
5. How many moles are there in 12g of CO2 (C=12, 0=16)?A. 0.27 B. 0.47 C. 0.16 D. 0.32

**THEORY**

1. Calculate the actual number of atoms contained in 2.8dm3 of chlorine (Molar volume of gas = 22.4dm3, NA = 6.02 X 1023)
2. How many moles are there in 10g of iron (II) tetraoxosulphate (VI)?

**WEEK TWO**

**TOPIC: WRITNG AND BALANCING CHEMICAL EQUATIONS**

Chemical equations are representation of chemical reactions in terms of the symbols and formulae of the elements and compounds involved. In a chemical equation, the reactants are always written on the left hand side while the products are written on the right hand side. For instance, if A and B combines together to give C and D, the equation of the reaction is written as:

A + B → C + D

Reactants Products

**BALANCING CHEMICAL EQUATIONS**

All equations must be balanced in order to comply with the law of conservation of matter. Equations are balanced through the use of coefficients in front of the formula and not by changing the subscript numbers within the formulae of the products.

Example 1: Write a balanced equation for the combustion of ammonia gas in air.

**Solution:**

Step I: Write the reactants and predict the products

NH3(g) + O2(g) → NO(g) + H2O(g)

Step II: The equation is not balanced. Therefore the equation can be balanced by placing the right coefficient in front of each molecule to balance the number of atoms. Thus, the balanced equation is:

4NH3(g) + 5O2(g) → 4NO(g) + 6H2O(g)

Example 2: Write a balanced equation for the combustion of ethane in oxygen.

**Solution:**

The general formula for the combustion of Alkanes is

CxHy + (x + y/4) O2 → XCO2 + y/2 H2O

The molecular formula for ethane is C2H6, so, x=2 and y=6

Substituting x and y into the formula above gives

C2H6 + (2 + 6/4) O2 →2CO2 + 6/2 H2O

C2H6 + 7/2 O2 →2CO2 + 3H2O

The equation is balanced. However, equations are written with whole number coefficients. By multiplying the entire equation by 2, we get

2C2H6 + 7O2 →4CO2 + 6H2O

**IMPORTANCE OF CHEMICAL EQUATIONS**

1. It gives us information on the product that can be formed from the combination of two or more reactants in a particular reaction.
2. It tells us the physical states of the reactants and products.
3. It indicates the direction of the reaction and whether the reaction is reversible.
4. It tells us the stoichiometry of the reaction (i.e. the relationship between the amount of reactants and products) in terms of mole ratio of the reactants and products involved.

Consider the table below:

|  |  |
| --- | --- |
| Equation | Mole ratio/ Mass ratio |
| 2HCl + CaCO3→ CaCl2 + H2O + CO2 | 2 mole of HCl and 1 mole of CaCO3produced 1 mole of CaCl2, 1 mole of H2Oand 1 moles of CO2 |
| 2HCl + CaCO3 → CaCl2 + H2O + CO2 | 73g of HCl and 100g of CaCO3produced 111g of CaCl2, 18g of H2Oand 44g of CO2 |

**GENERAL EVALUATION/REVISION**

1. Balance the following equations:

(a) KClO3(s) → KCl(s) + O2(g)

(b) ZnCO3(s) + HCl(aq) → ZnCl(aq) + H2O + CO2(g)

1. What is the volume in dm3 of 8g of oxygen gas at s.t.p?
2. State the use of each of the following apparatuses: triangular pipe clay, beehive shelf, bell jar, fume cupboard, dessicator.
3. Outline three differences between physical and chemical changes.

**WEEKEND ASSIGNMENT**

1. A balanced chemical equation obeys which of the laws? A. Law of conservation of matter B. Law of definite proportion C. Law of multiple proportion D. Boyle’s law
2. The numerical coefficients in a balanced equation give the A. number of mole of reactants and products B. molar mass of the reactants and products C. number of reactants only D. mass ratio of the reactants.
3. A molecule of neon is A. diatomic B. monoatomic C.triatomic D. polyatomic
4. H2SO4 + xKOH → K2SO4 + yH2O. The value for x and y in the above equation is A. 1 and 2 B. 2 and 3 C. 2 and 1 D. 4 and2
5. The balanced chemical equation for the reaction between hydrochloric acid and sodium hydroxide isA. NaOH + HCl → NaCl + H2O B. NaCl + HCl → NaOH + H2O

C. NaOH + H2SO4 →Na2SO4 + H2O D. H2SO4 + KOH → K2SO4 + H2O

**THEORY**

1. Balance the following equation: H2SO4 + Na2CO3 → Na2SO4 + H2O + CO2

    Ca(OH)2 + CO2 → CaCO3 + H2O

2. State two information provided by the equation of a chemical reaction.

**WEEK THREE**

**TOPIC: STOICHIOMETRY OF REACTIONS**

**CONTENT**

* Calculation of masses of reactants and products
* Calculation of volume of reacting gases

**STOICHIOMETRY OF REACTIONS**

The calculation of the amounts (generally measured in moles or grams) of reactants and products involved in a chemical reaction is known as stoichiometry of reaction. In other words, the mole ratio in which reactants combine and products are formed gives the stoichiometry of the reactions.

From the stoichiometry of a given balanced chemical equation, the mass or volume of the reactant needed for the reaction or products formed can be calculated.

**CALCULATION OF MASSES OF REACTANTS AND PRODUCTS**

1. Calculate the mass of solid product obtained when 16.8g of NaHCO3 was heated strongly     until there was no further change.

**Solution:**

    The equation for the reaction is:

    2NaHCO3(s) → Na2CO3(s) + H2O(g) CO2(g)

    Molar mass of NaHCO3 = 23 + 12 + 16x3 = 84gmol-1

    Molar mass of Na2CO3 = 23x2 +12+16x3 = 106gmol-1

    From the equation:

    2 moles NaHCO3 produces 1 mole Na2CO3

    2x84g NaHCO3 produces 106g Na2CO3

    16.8g NaHCO3 will produce Xg Na2CO3

     Xg Na2CO3 = 106g x 16.8g =10.6g

     2x84g

     Mass of solid product obtained = 10.6g

2. Calculate the number of moles of CaCl2 that can be obtained from 25g of limestone [CaCO3] in the presence of excess acid.

**Solution:**

      The equation for the reaction is:

       CaCO3(s) + 2HCl → CaCl2(s) + H20(l) + CO2(g)

       Number of moles = Reacting mass

                                 Molar mass

       Molar mass of CaCO3 = 40 + 12 + 16x3 = 100gmol-1

       Number of moles of CaCO3 = 25g = 0.25 mole

      100gmol-1

    From the equation of reaction,

    1 mole CaCO3 yields 1 mole CaCl2

    Therefore, 0.25 mole CaCO3 yielded 0.25 mole CaCl2.

**EVALUATION**

1. What does the term ‘Stoichiometry of reaction’ mean?
2. Ethane [C2H6] burns completely in oxygen. what amount in moles of CO2 will be produced when 6.0g of ethane are completely burnt in oxygen

**CALCULATION OF VOLUME OF REACTING GASES**

1. In an experiment, 10cm3 of ethene [C2H4] was burnt in 50cm3 of oxygen.

A. Which gas was supplied in excess? Calculate the volume of the excess gas remaining at the end of the reaction.

B. Calculate the volume of CO2 gas produced.

**Solution:**

The equation for the reaction is:

C2H4(g) + 3O2(g) → 2CO2(g) + 2H2O(g)

From the equation,

A.1 mole of ethene reacts with 3mole of oxygen

     1 volume of ethene reacts with 3 volumes of oxygen

    10cm3of ethene will react with 30cm3 of oxygen

    Since 50cm3 of oxygen was supplied, oxygen was in excess

    Hence volume of the excess gas = initial volume – volume used up = 50-30 = 20cm3

B.1 volume of ethene produces 2 volumes of CO2

    10 cm3 of ethene will produce 20cm3 of CO2

    Therefore, 20cm3 of CO2 was produced

2. 20cm3 of CO was mixed and sparked with 200cm3 of air containing 21% of O2. If all the     volumes are measured at s.t.p, calculate the total volume of the resulting gases.

**Solution:**

     In 200cm3 of air,

    Volume of O2 = 21 x 200cm3 = 42cm3

     100

Volume of N2 and rare gases = 200-42 = 158cm3

The equation for the reaction is:

2CO(g) + O2(g) → 2CO2(g)

Volume ratio 2 : 1 : 2

Before sparking 20cm3 42cm3

Reacting volume 20cm3 10cm3

After sparking 32cm3 20cm3

Volume of resulting gases = 32 + 20 + 158 = 210cm3

**GENERAL EVALUATION/REVISION**

1. Find the volume of oxygen produced by 1 mole of KClO3 at s.t.p in the following reaction:      2KClO3(s)  → 2KCl(s) + 302(g)

2. Balance the following equations: (a) Cu2S(s) +O2(g) → Cu2O(s) + SO2(g)

                                                  (b) C(s) + H2O(g) → CO(g) + H2(g)

3. Write the symbols of the following elements: mercury, silver, gold, lead, tin, antimony.

4. Define the term valency

**WEEKEND ASSIGNMENT**

1. Amount of a substance is expressed in A. mole B. grams C. kilograms D. mass
2. Determine the mass of CO2produced by burning 104g of ethyne [C2H2]A. 256g B.352g C. 416g D. 512g
3. The mole ratio in which reactants combine and products are formed is known as A. rate of reaction

B. stoichiometry of reaction C. equation of reaction D. chemical reaction

1. The unit for relative molecular mass is A. mole B. gmol-1C. grams D. mass
2. What mass of Pb(NO3) would be required to 9g of PbCl2 on the addition of excess NaCl solution? [Pb=207, Na=23, O=16, N=14] A. 10.7g B. 1.2g C. 6.4g D. 5.2g

**THEORY**

1. Calculate the number of molecules of CO2produced when 10g of CaCO3 is treated with 100cm3 of 0.20moldm-3 HCl.
2. Calculate the volume of nitrogen that will be producedat s.t.p from the decomposition of 9.60g ammonium dioxonitrate(iii), NH4NO2.

**WEEK FOUR**

**TOPIC: EMPIRICAL AND MOLECULAR FORMULAE**

Empirical formula is the formula which shows the simplest whole number ratios of atoms present in a compound while molecular formula is the formula which shows the actual number of each kind of atoms present in the molecule. The molecular formula of a compound is a whole number multiple of its empirical formulae.

**CALCULATIONS**

1. An organic compound on analysis yielded 2.04g carbon, 0.34g hydrogen and 2.73g oxygen.

A. Calculate the empirical formulA.

B. If the relative molecular mass of the compound is 60. Calculate its molecular formulA.

**Solution:**

    Elements C H O

    Reacting mass 2.04 0.34 2.73

   Mole ratio = Reacting mass = 2.04 : 0.34 : 2.73

     Atomic mass 12 1 16

= 0.17 : 0.34 : 0.17

  Dividing through by the 0.17 : 0.34 : 0.17

Smallest value 0.17 0.17 0.17

    Whole number ratio 1 : 2 : 1

      The empirical formula = CH2O

Relative molecular mass of the compound = 60

Let the molecular formula = (CH2O)n

(CH2O)n  = 60

(12 + 1x2 +16)n = 60

30n = 60

n = 60/30 = 2

Therefore, the molecular formula is (CH2O)2  = C2H4O2

Calculate the empirical formula of an organic compound containing 81.8% carbon and 18.2% hydrogen

**Solution:**

Element C H

% Composition by mass 81.8 18.2

Mole ratio = % by mass = 81.8 : 18.2

Atomic mass 12 1

= 6.82 : 18.2

Dividing through by the 6.82 : 18.2

smallest value 6.82 6.82

Whole number ratio 1 : 2.67

Since the ratio is not completely whole, we continue to multiple to obtain the lowest multiple that is close to a whole number i.e.

1:2.67, 2:5.34, 3:8.01, 4:10.65, 5:13.35, etC. 3:8.01 is close to whole number.

Therefore, the empirical formula is C3H8

**GENERAL EVALUATION/REVISION**

1. An organic compound has the empirical formula CH2. If its molecular mass is 42gmol-1,      what is the molecular formula?
2. Determine the relative molecular mass of calcium trioxocarbonate (v).
3. Define the term radical.
4. Write the formula of the following compounds

A. Mercury (i) dioxonitrate (iii)

B.Sodium hydrogen trioxocarbonate (IV)

C. Oxochlorate (I) acid

**WEEKEND ASSIGNMENT**

1. The % by mass of carbon in CO2 is A.37% B. 27% C. 48% D. 52%
2. What is the molar mass of Na2SO4? A. 172 B.168 C.142 D.133
3. The empirical formula of a compound is CH, the molecular formula could be A. C2H4

B.CH4C. C7H12D. C6H6

1. An oxide of nitrogen contains 69.6% of oxygen by mass. Its empirical formula is A. N2O3      B. N2O2C. N2O D. NO2
2. 5.0g of an oxide of a metal (M) gave 4.0g of the metal when reduced with hydrogen. What is the empirical formula of the oxide? [M=64, O=16] A. MO B. MO2C. M2O D. M203

**THEORY**

1. Calculate the % by mass of water of crystallization in Al2(SO4)3.9H2O
2. Two compounds X and Y have the same % composition by mass 92.3% carbon and 7.7% hydrogen. Calculate the:

A. Empirical formula of X and Y

B. Molecular formula of each compound if the molar mass of X is 26gmol-1 and Y is 78gmol-1.

**WEEK FIVE**

**TOPIC: LAWS OF CHEMICAL COMBINATION**

**CONTENT**

* Law of conservation of mass
* Law of definite proportion or constant composition
* Law of multiple proportion

**LAW OF CONSERVATION MASS**

This law states that during chemical reactions, matter can neither be created nor destroyed but changes from one form to another.

**EXPERIMENT TO VERIFY THE LAW**

AIM: To verify the law of conservation of mass

**THEORY**: The equation of the reaction chosen for study is as follows:

HCl(aq) + AgNO3(aq) → AgCl(s) + HNO3(aq)

White ppt

APPARATUS: Weighing balance, conical flask, small test tube, string cork stopper.

REAGENTS NEEDED: Solutions of HCl and AgNO3 stored in two different reagent bottles.

METHOD: The dilute HCl is poured into a conical flask. The small test tube is filled with  AgNO3solution and by means of a string tied around the neck of the test tube, it is  suspended inside the conical flask containing the acid in such a way that the two solutions  do not mix together. The conical flask and its content are weighed using a weighing  balance and the result recordeD. The two solutions are mixed together by swirling the  conical flask and the weight of the conical flask and its content is taken again.

DIAGRAM

RESULT: After mixing the two solutions, a white precipitate of AgCl was formed  indicating that a chemical reaction has taken place.

DISCUSSION: The masses of the conical flask and its content before and after the  reaction remained the same indicating that the mass of the reactants equal that of the   products.

CONCLUSION:Since the two masses obtained are equal, it confirms that matter was not created nor destroyed during the chemical reaction.

LAW OF DEFINITE PROPORTION OR CONSTANT COMPOSITION

The law states that all pure samples of a particular chemical compound contain the same    elements combined in the same proportion by mass.

EXPERIMENT TO VERIFY THE LAW

AIM: To verify the law of definite proportion

APPARATUS: Crucible, test tube, combustion boats, combustion tube, weighing balance,     Bunsen burner, U-tube and two retort stands with clamps.

REAGENTS NEEDED: CuCO3 crystals, Na2CO3 solution, Cu(NO3)2 solution, dry hydrogen     gas and CaCl2 crystals.

METHOD: Two samples of black CuO are prepared using different methods. Sample A is   prepared by placing the CuCO3 crystals in a crucible and heating it strongly until it   decomposes into black CuO. The equation for the reaction is:

CuCO3(s) → CuO(s) + CO2(g)

Sample B is prepared by reacting a solution of Na2CO3 in a test tube with a solution of  Cu(NO3)2. A green precipitate of CuCO3 is formeD. This is filtered off and then heated  strongly in a crucible to obtain black CuO. The equation for the reaction is:

Na2CO3(aq) + Cu(NO3)2 → CuCO3(s) + 2NaNO3(aq)

CuCO3(s) → CuO(s) + CO2(g)

The two samples of black CuO are placed in two dried and weighed combustion boats  labelled A and B and weighed again. These boats are then placed in a combustion tube and  heateD. A stream of dry hydrogen is passed through the combustion tube to reduce the  CuO to metallic Cu. After heating for sometimes, a reddish-brown residue shows that all  the CuO has been reduced to metallic copper. The flame is removed but the passing in  hydrogen gas continues to prevent the re-oxidation of the hot copper residues by  atmospheric oxygen. Any water formed during the reaction is absorbed by the fused  CaCl2 in the adjacent U-tube. When the boat is cool, the weight of it is taken. From the  results, the percentage of Cu in each sample is calculated.

DIAGRAM:

RESULT: Assuming the following result was obtained:

Sample A B

Mass of boat 3.16g 3.31g

Mass of boat + CuO 5.15g 5.29g

Mass of boat + Cu 4.76g 4.90g

Mass of CuO = (ii) – (i) 1.99g 1.98g

Mass of Cu = (iii) –(i) 1.60g 1.59g

% of Cu in CuO 1.60 x 100 1.59 x 100

1.99 1 1.98 1

80.40% 80.30%

Therefore, % of Cu in CuO 80% 80%

% of O2 in CuO 20% 20%

DISCUSSION: The % of Cu residue in the two samples is approximately 80% irrespective of the method of preparation of the CuO samples.

CONCLUSION: In pure CuO, Cu and O are always present in a definite proportion by mass of approximately 4:1.

LAW OF MULTIPLE PROPORTIONS

This states that if two elements combine to form more than one compound, the masses of one of the elements which separately combine with fixed mass of the other element are in simple ratio

EXPERIMENT TO VERIFY THE LAW

AIM: To verify the law of multiple proportions

APPARATUS: Combustion boats, combustion tube, weighing balance, Bunsen burner, U-tube and retort stand with clamp

REAGENTS NEEDED: Cu2O crystals, CuO crystals, dry hydrogen gas and calcium chloride crystals

METHOD: The two boats are dried and weighed. Cu2O is placed in one and labelled A and CuO is placed in the other and labelled B. The two boats are weighed again and placed in a combustion tube to reduce the oxides to copper by passing hydrogen gas into the combustion tube. When the samples are cooled, the residues obtained are weighed.

RESULT: Assuming the following result was obtained:

Sample Cu2O CuO

Mass of sample (oxide) 3.04g 1.91g

Mass of Cu residue 2.55g 1.35g

Mass of oxygen removed from oxide 0.49g 0.53g

CALCULATION: Calculating the various masses of copper which combine separately with fixed mass (say 1g of oxygen)

For Cu2O,

0.49g of O2 combines with 2.55g of Cu

1.0g of O2 will combine with Xg of Cu

Xg of Cu = 2.55g x 1.0g

0.49g

= 5.20g

For CuO,

0.53g of O2 combines with 1.38g of Cu

1.0g of O2 will combine with Xg of Cu

Xg of Cu = 1.38g x 1.0g

0.53g

= 2.60g

Oxides of copper Cu2O CuO

Mass of copper 5.20g 2.60g

Ratio of copper 2 : 1

CONCLUSION: The masses of copper which combines with a fixed mass of oxygen in Cu2O and CuO are in simple ratio of 2:1.

**GENERAL EVALUATION/REVISION**

1. State the law of (a) definite proportion (b) multiple proportion
2. Balance the following chemical equation
3. Ca(HCO3)2(s) → CaCO3(s) + H2O(l) + CO2(g)
4. SO2(g) + H2O(l) + 02(g) → H2SO4(aq)
5. Determine the oxidation number of: (a) Cu in CuCO3 (b) P in H3PO4 and name the compound

**WEEKEND ASSIGNMENT**

1. All pure samples of chemical compound contain the same element in the same proportion    by mass is a the law of------ A. definite proportion B. reciprocal proportion C. multiple    proportion

D. conservation of matter

1. What is used to measure the mass of atom and molecules? A. Beam balance B. Spring balance

C. Chemical balance   D. Mass spectrometer

1. What is the ratio by mass of oxygen and hydrogen in 1 mole of water?A. 3:1 B. 2:1 C.1:2 D. 2:4
2. In two separate experiments 0.18g and 0.36g of chlorine combine with a metal M, to give A and B respectively. An analysis showed that A and B contain 0.10g and 0.20g of M     respectively. Which law is illustrated by the data? A A. Law of multiple proportions. B.Law of conservation of mass.

C. Law of constant composition. D. Law of simple proportion

1. An element E forms the following compounds with bromine: EBr2, EBr3, and EBr4. This observation illustrates the A. Law of conservation of mass.B. Law of definite    proportion.C. Law of multiple proportion.D. Law of chemical combination

**THEORY**

1. 8.50g of CuO when heated in a current of dry hydrogen gas gave 6.58g of copper and 2.16g of water. Calculate the proportion of oxygen to hydrogen by mass in water.
2. Balance the following equations:

C4H10 + O2 → CO2 + H2O

H2SO4 + NaOH → Na2SO4 + H2O

**WEEK SIX AND SEVEN DATE-----------------**

**TOPIC: CHEMICAL COMBINATIONS**

**CONTENT**

* Electrovalent (ionic) bond
* Covalent bond
* Dative bond
* Hydrogen bond
* Metallic bond

**ELECTROVALENT (IONIC) BOND**

Electrovalent bond is characterised by transfer of electrons from metallic atoms to non-metallic atoms during reaction. The metallic atom that donates electron becomes positively charged while the non-metallic atom that accepts electron becomes negatively charged. The strong electrostatic attraction that holds the oppositely charged ions together is called ionic bond.

**ELECTRON DOT REPRESENTATION OF FORMATION OF IONIC COMPOUNDS**

Formation of sodium chloride

Formation of calcium oxide

**PROPERTIES OF SOME IONIC COMPOUNDS**

1. They are solids at room temperature.
2. They contain oppositely charged ions.
3. They readily dissolve in water and other polar solvents like ethanol.
4. They have high melting and boiling points
5. They are good conductors of electricity when in molten or in aqueous form.

**EVALUATION**

1. How is an ionic compound formed?
2. State the properties of ionic compound

**COVALENT BOND**

This involves the sharing of a paired of electron between two reacting atoms. The shared electrons are each contributed by the reacting atoms and are called shared pair. A shared pair of electron in covalent bond is represented by a horizontal line

(----) between the two atoms

**ELECTRON DOT REPRESENTATION OF FORMATION OF COVALENT COMPOUNDS**

Formation of hydrogen molecule

Formation of carbon (iv) oxide

**PROPERTIES OF COVALENT COMPOUNDS**

1. They consist of molecules with definite shape.
2. They are gases or volatile liquids.
3. They readily dissolve is non-polar organic solvents
4. They have low melting and boiling points
5. They do not conduct electricity because the molecules do not contain charged particles.

**EVALUATION**

1. What is covalent bond?
2. Outline the properties of covalent compounds

**COORDINATE COVALENT (DATIVE) BOND**

In coordinate covalent bond, the shared pair of electrons is supplied by one of the combining atoms. Coordinate covalent bond is often formed in molecules that have a lone pair of electrons, i.e. a pair of electron not directly concerned in an existing bond.

**ELECTRON DOT REPRESENTATION TO SHOW FORMATION OF DATIVE BOND**

Formation of hydroxonium ion (H3O+)

Formation of Ammonium ion (NH4+)

Compounds containing coordinate covalent bond are similar in properties to purely covalent compounds. Both do not conduct electricity, but the presence of coordinate covalent bond tends to make a compound less volatile.

**HYDROGEN BOND**

Hydrogen bond is a dipole-dipole intermolecular force of attraction which exists when hydrogen is covalently bonded to a highly electronegative element of small atomic size. The electronegative element can be N, O, F, Cl, Br or I.

The highly electronegative element has very strong affinity for electrons. Hence, they attract the shared pair of electrons in the covalent bond toward themselves, resulting in the formation of a dipole which leaves a partial positive charge on the hydrogen atom and a partial negative charge on the electronegative atom. An electrostatic attraction between two dipoles is set up when the positive pole of one molecule attracts the negative pole of the other. This attractive force is known as hydrogen bond.

**IMPORTANCE OF HYDROGEN BOND**

It accounts for the solubilities of some compounds containing O, N and F in certain hydrogen containing solvents such as water

The crystalline shape of solid water (ice) is due to hydrogen bond.

**EVALUATION**

1. Define hydrogen bond

2. State two importance of hydrogen bond.

**METALLIC BOND**

Metal atoms are held together in solid crystal lattice by metallic bond. each metallic atom contributes its outer (valence) electron to the electron cloud, thus becoming positively charged. The resulting positively charged metallic ions tend to repel each other but are held together by the moving electron cloud and overlapping residual electron orbits. Thus, a metallic bond is a force of attraction between the positive metal ions and the free mobile electrons.

**VAN DER WAALS’ FORCES**

The attractive forces which make it possible for non-polar molecules like nitrogen and CO­2 molecules to form liquid and solid is called van der Waals’ force. This force though very weak when compared to ionic and covalent bond is important in the liquefaction of gases and in the formation of molecular lattices as in iodine and naphthalene crystals.

**GENERAL EVALUATION/REVISION**

1. Using electron dot representation, show the formation of MgO and O2 molecule
2. Define hydrogen bond
3. How is metallic bond formed?
4. Describe how you will separate a mixture of NaCl, Iodine and PbCl2

**WEEKEND ASSIGNMENT**

1. Noble gases are stable because they A. are volatile B. have octet configuration C. have no neutron in their nucleusD. forms ions easily
2. The bond type in diatomic nitrogen gas is A. double covalent bondB. triple covalent bondC. single covalent bondD. double electrovalent bond
3. In electrovalency, valence electrons are transferred and the atomic number is A. reduced B. stabilized C. unaffected D. increased
4. An element Y having an atomic number of 19 combines with another element Z with     atomic number 17. The likely compound formed is A. Y2Z B. Y3Z2C. YZ D. Y2Z2
5. The type of attractive force which exist between discrete molecules is calledA. metallic bond

B. hydrogen bondC. dative bondD.van der Waals’ forces

**THEORY**

1. A. Illustrate the formation of the compound AlCl3 using electron dot representation

B. State two properties of the compound

1. Define hydrogen bond

**WEEK EIGHT DATE-----------------**

**TOPIC: KINETIC THEORY OF MATTER**

**CONTENT**

* States of matter
* Change of state: melting, boiling, evaporation, condensation and freezing
* Kinetic theory of gases
* Phemomena supporting kinetic theory of matter

**STATES OF MATTER**

The three states of matter: solid, liquid and gaseous states can be distinguished by the motion of particles they are made of and the attractive force between their particles.

|  |  |  |
| --- | --- | --- |
| **SOLID** | **LIQUID** | **GASES** |
| Have definite shape and volume | Have no definite shape but definite volume | Have no definite shape and volume |
| Very dense | Less dense | Least dense |
| Incompressible | Incompressible | Compressible |
| Fixed mass | Fixed mass | Fixed mass |
| Particle vibrate and rotate about a fixed point | Particles vibrate and move about within a restricted space | Particles move about constantly at great speed and at random |

**CHANGE OF STATE**

**MELTING**

Melting is the physical process where a substance changes from a solid to a liquid. When a solid is heated, the particles acquire greater kinetic energy and move violently. A point is reached when the forces of vibration overcome the cohesive forces holding the solid particles together and the crystalline structure collapses. The particles are no longer held in fixed positions but are free to move about and the liquid state is reached. The temperature at which this occurs is called the melting point of the solid.

**BOILING**

When a liquid is heated, the rate of evaporation increases and the value of the saturated vapour pressure equal the prevailing atmospheric pressure. When this happens, the liquid is said to boil and the temperature at which this happen is known as the boiling point of the liquid.

The boiling point of a liquid change with change in atmospheric pressure. If the pressure is raised, the boiling point will increase and if the pressure is lowered the boiling point will decrease. Also, the presence of impurities increases the boiling point of a liquid.

**EVAPORATION**

Evaporation is the process of vapourization of liquids at all temperatures. When the surface of a liquid is exposed, the molecules near the surface of the liquid will acquire extra kinetic energy, large enough to enable them break away from the cohesive force binding them to the neighbouring particles. Once free, they escape from the liquid surface to become molecules in the vapour state.

Evaporation results in decrease in the volume of liquid and lowering the temperature of the liquid, therefore it causes cooling. Also, it occurs at all temperature but increases with increase in temperature. In addition, it is slower in electrovalent liquids than in covalent liquids.

**DIFFERENCES BETWWEEN EVAPORATION AND BOILING**

|  |  |
| --- | --- |
| **EVAPORATION** | **BOILING** |
| Takes place at the surface of the liquid | Involves the entire volume of the liquid |
| Takes place at all temperature | Takes place at a fixed temperature |

**CONDENSATION AND FREEZING**

Condensation is a process whereby a vapour loses some of its kinetic energy to a colder body and changes into the liquid state.

When a liquid cools, it loses heat energy to its surroundings, causing its temperature to drop. If the cooling continues, the temperature of the liquid keeps dropping until it reaches the freezing point of the liquid. At this temperature, the liquid changes into solid.

**EVALUATION**

1. Describe the melting process of a solid.
2. State two differences between evaporation and boiling.

**KINETIC THEORY OF GASES**

The theorypostulates the following for an ideal or perfect gas:

Gas molecules are in constant, rapid, straight motion, colliding with one another and with the walls of the container.

The collision of gas molecules is perfectly elastic.

The total volume of the gas molecule is negligible compared to the volume of the container.

The force of attraction between the gas molecules is negligible.

The average kinetic energy of the molecule is a measure of the temperature of the gas molecules.

**PHENOMENA SUPPORTING THE KINETIC THEORY OF GASES**

Brownian motion: This is the constant, irregular movement of particles in a liquid or gas. It shows that gas molecules are in constant motion.

Diffusion: Diffusion is the movement of particles from a region of higher concentration to lower concentration. Diffusion is common in gases and it results from the random movement of particles of a gas.

**GENERAL EVALUATION/REVISION**

1. Compare the three states of matter under the following headings: Shape/volume,    Density, Compressibility and Motion of particles.
2. Write short note on (a) Boiling (b) Evaporation.
3. 100cm3 each of 0.02moldm-3 solution of HCl and Pb(NO3)2 were mixed. Assuming the     PbCl2 is completely insoluble; determine the mass of the PbCl2 precipitated.
4. State the postulates of Dalton’s Atomic theory.

**WEEKEND ASSIGNMENT**

1. ------ is measure of the average kinetic energy of the molecules of a gas. A. mass B. volume

C. pressure D. temperature.

1. All the following are the assumptions of the kinetic theoryof gases except A. Gases are composed of many elastic particles called molecules. B. The molecules are of negligible C. The molecules collide with one another and with the walls of container.D. The molecules are in constant random motion.
2. Presence of sodium chloride in ice will A. decrease the melting point of the ice B. increase the melting point of the ice C. make sodium chloride impure D. lower the freezing point of sodium chloride
3. Which of these does not support the kinetic theory? A. Brownian motion B. Diffusion C.Osmosis

D. Linear expansivity

1. The phenomenon whereby the atmospheric pressure equals the saturated vapour pressure is called

A. freezing B. latent heat C. boiling D. normal pressure

**THEORY**

1. A bottle of milk is taken out of the refrigerator and placed on the table. Droplets of     water are noticed on the surface of the milk bottle. Explain the observation
2. State two phenomena that support the kinetic theoryof gases.

**WEEK NINE**

**TOPIC: GAS LAWS**

**CONTENT**

* Boyle’s law
* Charles’ law
* Ideal gas equation
* Dalton’s law of partial pressure

**BOYLE’S LAW**

It states that the volume of a fixed mass of gas is inversely proportional to the pressure provided the temperature remains constant.

Mathematically,

V α 1/P

V = k/P

PV = k

Hence, P1V1 = P2V2

Boyle’s law can be represented graphically as shown below.

The graph shows that if the pressure is doubled, the volume is reduced to half its former value and if it is halved, the volume is doubled.

**EXPLANATION OF BOYLE’S LAW USING THE KINETIC THEORY**

When the volume of fixed mass of gas is decreased, the molecules of the gas will collide with each other more rapidly. This gives rise to an increase in pressure. However, if molecules are farther apart the number of collisions for unit time decreases, resulting in a decrease in pressure.

**CHARLES’ LAW**

Charles’ law states that the volume of a fixed mass of gas at constant pressure is directly proportional to its temperature in the Kelvin scale.

Mathematically,

V α T

V = k/T

V = k

T

Hence, V1 = V2

T1 T2

The graphical representation of Charles’ law is as shown below:

**EXPLANATION OF CHARLES’ LAW USING THE KINETIC THEORY**

When a given gas is heated at constant pressure, the molecules acquire more kinetic energy and move faster. They collide with one another and with the walls of the container more frequently. To maintain the same number of collisions on the walls of container (i.e. keep the pressure constant) the volume of the gas increases.

**CALCULATIONS BASED ON BOYLE’S AND CHARLES’ LAW**

1. 200cm3 of a gas has a pressure of 510mmHg. What will be its volume if pressure in increased to 780mmHg, assuming there is no change in temperature?

**Solution:**

    V1 = 200cm3, P1 = 510mmHg, P2 = 780mmHg V2 = ?

    Using the expression for Boyle’s law:

P1V1 = P2V2

V2 = P1V1 = 510mmHg x 200cm3 = 130.769 = 131 cm3

P2 780mmHg

2. A certain mass of a gas occupies 300cm3 at 35oC. At what temperature will it have its volume reduced by half assuming its pressure remains constant?

**Solution:**

       V1 = 300cm3, T1 = 35oC = (35 + 273)K = 308K, V2 = V1/2 = 300/2 = 150cm3, T2 = ?

       Using the formula for Charles’ law

V1 = V2

T1 T2

T2 = V2T1 = 150cm3 x 308K = 154K

V1 300cm3

**EVALUATION**

1. State Boyle’s law

2. Explain Charles’ law using the kinetic theory

**GENERAL GAS EQUATION**

Boyle’s and Charles’ laws are combined into a single expression known as the general gas equation which can be expressed mathematically as

P1V1 = P2V2

T1T2

**IDEAL GAS EQUATION**

This equation states that for an ideal gas PV/T is a constant.

That is, PV = R (R = molar gas constant)

T

PV = RT

That is, for n mole of a gas, the equation becomes

PV = nRT

**CALCULATIONS**

1. What is the volume at s.t.p of a fixed mass of a gas that occupies 700cm3 at 25oC and 0.84 x 105 Nm-2pressure?

**Solution:**

   T1 = 273K, P1 = 1.01 x 105Nm-2, T2 = 25oC = (25 + 273) = 298K, P2 = 0.84 x 105Nm-2,

    V2 = 700cm3, V1 =?

   Using the general gas equation

P1V1 = P2V2

T1 T2

V1 = P2V2T1 = 0.84 x 105Nm-2 x 700cm3 x 273K = 533.337 =533cm3

P1T2 1.01 x 105Nm-2 x 298K

2. Calculate the number of moles present in a certain mass of gas occupying 6.5dm3 at     3atm and 15oC (R = 0.082atmdm3K-1mol-1)

**Solution:**

    V = 6.5dm3, P = 3atm, T = 15oC = (15 + 273)K = 288K, n =?

    Using PV = nRT

n = PV = 3atm x 6.5dm3= 0.8257

RT 0.082atmdm3K-1mol-1 x 288K

    Number of moles = 0.83 mole

**DALTON’S LAW OF PARTIAL PRESSURE**

This law state that in a mixture of gases which do not react chemically together, the total pressure exerted by the mixture of gases is equal to the sum of the partial pressure of the individual gases that make up the mixture.

Mathematically, the law can be expressed as:

Ptotal = PA + PB +PC........Pn

Where Ptotal is the total pressure of the mixture and PA, PB, PC are the partial pressure exerted separately by the individual gases A, B, C that make up the mixture.

The pressure each constituent gas exerts is called partial pressure and is expressed as

Partial pressure of gas A (PA) = Number of moles of gas A x Ptotal

Total number of moles of gas in mixture

That is, PA = nA x Ptotal

nA + nB + nC

If the gas is collected over water, it is likely to be saturated with water vapour and the total pressure becomes

Ptotal = Pgas + Pwater vapour

Pgas = Ptotal – Pwater vapour

**CALCULATION ON THE LAW**

A gaseous mixture containing 64g of O2 and 70g of N2 exerts a total pressure of 1.8oatm. What is the partial pressure exerted by oxygen in the mixture?

**Solution:**

Molar mass of O2 = 16 x 2 = 32gmol-1

Molar mass of N2 = 14 x 2 = 28gmol-1

Number of mole of O2 = 64g = 2.0mole

                          32gmol-1

Number of mole of O2 = 70g = 2.5mole

                          28gmol-1

Total number of moles of gases in mixture = 2.0 + 2.5 = 4.5 mole

Partial pressure of O2 = 2.0 x 1.80 = 0.80atm

                        4.5

**GENERAL EVALUATION/REVISION**

1. State Dalton’s of partial pressure.
2. Calculate the pressure at 27oC of 16.0g O2 gas occupying 2.50dm3
3. A certain mass of hydrogen gas collected over water at 10oC and 760mmHg pressure has     a volume of 37cm3. Calculate the volume when it is dry at s.t.p (Saturated vapour     pressure of water at 10oC =1.2mmHg)
4. Determine the number of electrons, protons and neutrons in each of the following: 39K19,63.5Cu29.
5. If an element R has isotopes 60% of 12R6 and 40% xR6 and the relative atomic mass is      12.4, find x.

**WEEKEND ASSIGNMENT**

1. Kelvin temperature can be converted into temperature by A.oC = K-273 B. K + 273 C.oC + 273/K D. K + 273/oC
2. The pressure exerted by a gas is a result of the A. continuous random motion of its particle.

B. bombardment of the walls of the container by its molecules. C.expansion of the gas molecules

D. collision between the gas molecules.

1. From the ideal gas equation, PV = nRT, the unit of n is A. atmdm3B. atmdm3/K C. mole D. K/mole
2. What will be the new volume (V) if the new pressure is halved and the initial pressure      remain the same the sameA. 2P1V1 = P2V2B. P1V1 = 2P2V2C. P1V1/2= P2V2/2 D. P1V1 = P2V2/2
3. A fixed mass of gas of volume 546cm3 at 0oC is heated at constant pressure. What is     the volume of the gas at 2oC? A. 550cm3B. 560cm3C. 570cm3D. 580cm3

**THEORY**

1. A given mass of nitrogen is 0.12dm3 at 60oC and 1.01 x 105Nm-2. Find its pressure at the    same temperature if its volume is changed to 0.24dm3
2. 272cm3 of CO2 was collected over water at 15oC and 782mmHg pressure. Calculate the     volume of dry gas at s.t.p (saturated vapour pressure of water at 15oC is 12mmHg).

**WEEK TEN**

**TOPIC: GAS LAWS**

**CONTENT**

* Avogadro’s law
* Gay lussac’s law of combinig volumes
* Graham’s law of diffusion

**AVOGADRO’S LAW**

This law states that equal volume of all gases at the same temperature and pressure contain the same number of molecules. This means that 1 mole of any gas at s.t.p has a volume of 22.4dm3.

**GAY LUSSAC’S LAW OF COMBINING VOLUMES**

It states that when gases react they do so in volumes which are simple ratios to one another and to the volumes of the products if gaseous, provided that the temperature and pressure remain constant.

**CALCULATION ON THE LAW**

Calculate the volume of oxygen required to burn 500cm3 of methane completely.

**Solution:**

 The equation for the reaction is:

2CH4(g) + 3O2(g) → 2CO2(g) + 2H2O(g)

 By Gay Lussac’s law,

 2 volumes of CH4 requires 3 volumes of O2 for complete combustion

 Therefore, 2cm3 of CH4 requires 3cm3 of O2

500cm3 of CH4 will require Xcm3 of O2

  Xcm3 of O2 = 500cm3 x 3cm3 = 750cm3

2cm3

   Thus, 750cm3 of O2 is required

**EVALUATION**

1. State the Gay Lussac’s law of combining volumes
2. 40cm3 of hydrogen was sparked with 160cm3 of oxygen at 100oC and 1atm. Determine the volume of oxygen left after the reaction.

**GRAHAM’S LAW OF DIFFUSION**

It states that the rate of diffusion of a gas is inversely proportional to the square root of its density at constant temperature and pressure.

Mathematically,

R α 1/√d

R = k/√d where k is a constant

Comparing the rate of diffusion of two gases:

R1 = √d2

R2 √d1

In terms of relative molecular mass, M

R α 1/√M

For two gases,

R1 = √M2

R2 √M1

But rate of diffusion is reciprocal of time, R =1/t

That is,

R1 = t2

R2 t1

From the inverse relationship we can deduce that the less dense a gas is, the higher the rate of diffusion and vice versa.

**CALCULATION**

1. A given volume of SO2 diffuses in 60 seconds. How long will it take the same volume of    CH4 to diffuse under the same condition (SO2 = 64, CH4 = 16)

**Solution:**

  Using the expression:

t1= √M2

t2 √M1

t2 = √M2 x t1 = √16 x 60seconds = 30seconds

√M1  √64

**GENERAL EVALUATION/REVISION**

1. State Graham’s law of diffusion
2. Under the same condition of temperature and pressure, hydrogen diffuses 8 times as     fast as gas Y. Calculate the relative molecular mass of Y.
3. State the following rule/principle: (a) Hund’s rule of maximum multiplicity (b) Aufbau principle
4. Write the electronic configuration of (a) oxide ion, (b) Aluminium ion, (c) potassium (d) phosphorus.

**WEEKENDASSIGNMENT**

1. 400cm3 of a gas X diffuses through a porous pot in 2 minutes. Calculate the rate at which X diffuses. A. 6.3cm3s-1B. 20cm3s-1 C. 200cm3s-1D. 3.33cm3s-1#
2. The relationship between the density (d) of a gas and the rate at which the gas diffuses is A. R = kd

B. R= k/√d C. R = k√d D. k/d

1. Calculate the minimum volume of oxygen required for the complete combustion of a mixture of 20cm3 CO and 20cm3 of H2. A. 10cm3B. 20cm3C. 40cm3D. 60cm3
2. If sulphur (iv) oxide and methane (CH4) are released at the same time at opposite ends of a tube, the rate of diffusion will be in the ratio A. 2:1 B. 4:1 C. 1:4 D. 1:2
3. ‘Equal volume of all gases at the same temperature and pressure contain same number of molecules’ is a state of which law A. Avogadro’s law B. Boyle’s lawC. Charles’ law      D. Chemical law

**THEORY**

1. Arrange the following gases in order of increasing rate of diffusion: CO, SO2, H2S,NO2 and O2.
2. The vapour densities of O2 and Cl2 are 16 and 36 respectively. If 60cm3 of O2 diffuses through a porous partition in 14 seconds, how long will it take 1000cm3 of Cl2 to diffuse through the same partition?