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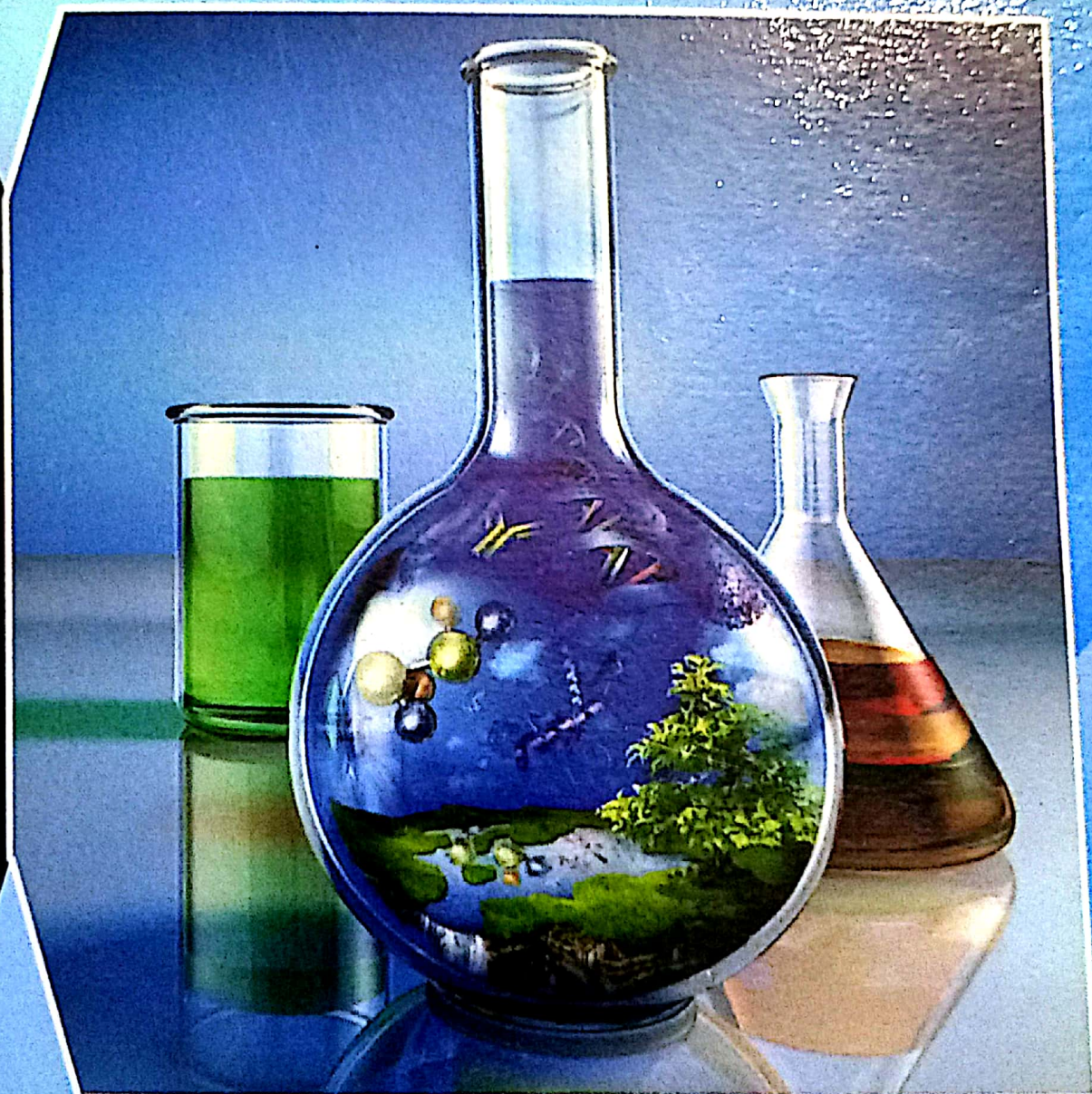
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**SOLVED**

Smart Syllabus 2020-2021

# CHEMIST

CHAPTER WISE SOLUTION  
BOARD SESSION 2012-2019



Intermediate Part - I

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**M.S.C Chemistry**

**Govt. Post Graduate College, Sahiwal**



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# CONTENTS

	ALP SMART SYLLABUS	3
CHAPTER-1	OBJECTIVE (MCQ'S)	5
	SHORT QUESTIONS	6
	SOLVED EXERCISE	14
CHAPTER-2	OBJECTIVE (MCQ'S)	25
	SHORT QUESTIONS	25
	SOLVED EXERCISE	27
CHAPTER-3	OBJECTIVE (MCQ'S)	29
	SHORT QUESTIONS	29
	SOLVED EXERCISE	39
CHAPTER-4	OBJECTIVE (MCQ'S)	45
	SHORT QUESTIONS	46
	SOLVED EXERCISE	52
CHAPTER-5	OBJECTIVE (MCQ'S)	57
	SHORT QUESTIONS	58
	SOLVED EXERCISE	65
CHAPTER-6	OBJECTIVE (MCQ'S)	71
	SHORT QUESTIONS	73
	SOLVED EXERCISE	84
CHAPTER-7	OBJECTIVE (MCQ'S)	87
	SHORT QUESTIONS	87
	SOLVED EXERCISE	92
CHAPTER-8	OBJECTIVE (MCQ'S)	97
	SHORT QUESTIONS	98
	SOLVED EXERCISE	110
CHAPTER-9	OBJECTIVE (MCQ'S)	118
	SHORT QUESTIONS	118
	SOLVED EXERCISE	124
CHAPTER-10	OBJECTIVE (MCQ'S)	128
	SHORT QUESTIONS	129
	SOLVED EXERCISE	139
CHAPTER-11	OBJECTIVE (MCQ'S)	144
	SHORT QUESTIONS	144
	SOLVED EXERCISE	148

## ALP SMART SYLLABUS 2020

### CHAPTER 1: BASIC CONCEPTS

TOPIC: (1.3(1.3.1, 1.3.3), 1.5, 1.6, 1.7, 1.8) Isotopes (Relative Abundance of Isotopes (Pg.3-4), Average Atomic Mass(Pg. 6-6), Concept of Mole, Stoichiometry, Limiting Reactant, Yield (Pg. 11-22)

Classwork: Q.1 (i, ii, iii, v, x), 2Q.(i, ii, iii, v, vii, viii), Q.3 (i, v, vi, viii) Q.9 to Q.18, Q.20, Q.21, Q.22, Q.25

Homework: Q.5 (a, b), Q.6, Q.7, Q.8 (vi, vii, viii)

### CHAPTER 2: EXPERIMENTAL TECHNIQUES IN CHEMISTRY

TOPIC: (2.3, 2.4, 2.5) Solvent-Extraction, Chromatography (Pg. 34 to 37).

Classwork: Q.1 (iii, iv, v) Q.2 (1, 4, 5) Q.3 (iv, v) Q.7

Homework: Q.6, Q.8, Q.9, Q.10

### CHAPTER 3: GASES

TOPIC: (3.2, 3.3, 3.4, 3.5, 3.7, 3.8, 3.11) Gas Laws, Dalton's Law of Partial Pressure (Pg. 41-57). Kinetic Molecular Theory of Gases, Kinetic Interpretation of Temperature (Pg. 60 - 65). Plasma State (Pg. 73 - 75).

Classwork: Q.1 (i, ii, iii, iv, v, vi, vii, viii), Q.2 (i, ii, iii, v) Q.3(i, ii), Q.8,9, 16,17, 18, 19, 20, 22, 23

Homework: Q.4, Q.5, Q.6, Q.7, Q.10, Q.12.

### CHAPTER 4: LIQUIDS AND SOLIDS

TOPIC: (4.1, 4.3, 4.4, 4.5, 4.6) Intermolecular Forces (Pg. 81 - 88). Crystal lattice, Crystals and Their Classification (Pg. 95 - 101). QUESTIONS ON LIQUIDS

Classwork: Q.1 (i, ii, iii, iv), Q.2 (i, ii, iii, iv, v, vi, viii), Q.3 (i, ii, iii, iv, v, vi, ix), Q.4, Q.5, Q.6, Q.7, Q.8.

Homework: Q.12. QUESTIONS ON SOLIDS

Classwork: Q.1 (ii, iii, iv) Q.2, Q.12(vi, vii, viii, ix, x, xi)

Homework: Q.4, Q.5, Q.6.

### CHAPTER NO 5: ATOMIC STRUCTURE

TOPIC: (5.1, 5.2, 5.3, 5.4, 5.5, 5.7, 5.8) Sub-Atomic Particles of Atoms, Rutherford's Model of Atom (Discovery of Nucleus), Plank's Quantum Theory, Bohr's Model of Atom, Spectrum (Pg. 118- 137). Wave-Particle Nature of Matter (Dual Nature of Matter), Heisenberg's Uncertainty Principle, (Pg. 138- 146).

Classwork: Q.1 (i, ii, iii, iv, v, vii, viii, ix, x), Q.2 (i to viii), Q.3, Q.4, Q.17, Q.19, Q.23, Q.24, Q.25

Homework: Q.5, Q.6, Q.7, Q.8, Q.9, Q.10, Q.11, Q.14, Q.15, Q.16.

### CHAPTER NO 6: CHEMICAL BONDING

TOPIC: (6.1, 6.2, 6.3, 6.4) Chemical Bond, Atomic Sizes, Ionization Energy, Electron Affinity and Electronegativity, Types of Bonds (Pg. 155 -182).

Classwork: Q.1 (i, ii, iii, v, vi), Q.2 (i, ii, iii, iv, v), Q.3 (i, ii, iii, iv, v, vii, viii, ix, x, xi, xii), Q.6, Q.10, Q.18 (ii, vi).

Homework: Q.4, Q.5, Q.7, Q.8, Q.9, Q.11, Q.12.

### CHAPTER NO: 7 THERMOCHEMISTRY

TOPIC: (7.2,7.3, 7.4, 7.5) System, Surrounding And State function, Internal Energy and First Law of Thermodynamics, Enthalpy, Hess's Law of Constant Heat Summation (197- 209).

Classwork: Q.1, Q.2 (i, ii, iii, v), Q.3(ii, iii, iv, v), Q.13, Q.14, Q.15, Q.16, Q.17, Q.18, Q.19, Q.20, Q.21.

Homework: Q.4, Q.5, Q.7, Q.8, Q.9, Q.10, Q.11, Q.12.

### CHAPTER NO: 8 CHEMICAL EQUILIBRIUM

TOPIC: (8.1, 8.2, 8.3, 8.4, 8.5,8.7, 8.8) Reversible and Irreversible Reactions, Application of Chemical Equilibrium In Industry, Ionic Product of water, Ionization Constants of Acids (K<sub>a</sub>), Ionization Constant of Bases (K<sub>b</sub>). (Pg. 214-235). Common Ion Effect, Buffer Solutions (Pg. 236-242).

Classwork: Q.1 (i, ii, iii, v), Q.2, Q.3 (i, ii, iii, iv), Q.10(a, b), Q.11, Q.19, Q.20, Q.21, Q.22, Q.23.

Homework: Q.6, Q.7, Q.8, Q.17.

### CHAPTER NO: 9 SOLUTION

TOPIC: (9.3, 9.5, 9.6, 9.7) Ideal and Non- Ideal Solutions (Pg. 260-262). Solubility and Solubility Curves, Colligative Properties of Solutions, Energetics of Solution (265-277).

Classwork: Q.1(ii, v, vi, vii, viii, ix, x), Q.2 (ii, iii, iv, v, vi, vii, viii, ix, x), Q.3(iii, iv, v, vi, vii, viii), Q.4, Q.5, Q.12, Q.12, Q.21, Q.22, Q.23.

Homework: Q.7, Q.8, Q.9, Q.10, Q.11, Q.13, Q.14, Q.15, Q.16.

### CHAPTER NO: 10 ELECTROCHEMISTRY

TOPIC: (10.1(10.1.1, 10.1.2), 10.2, 10.3, 10.4) Definition of Electrochemistry, Oxidation State and Balancing of Redox Equations (Oxidation Number or State, Finding Oxidation Number of an



Element in a compound or a Radical) (Pg. 284-285), Electrolytic Conduction, Electrode Potential, Electrochemical Series (Pg. 289-300).  
 Classwork: Q.1, Q.2 (i, ii, iii, iv, v), Q.3 (i, ii, iii, iv, v, vi, vii, viii), Q.4, Q.15, Q.16 (b, d, e, g, h).  
 Homework: Q.7, Q.8, Q.9, Q.10, Q.11, Q.12, Q.13, Q.14 (a, b).  
**CHAPTER NO: 11 REACTION KINETICS**  
**TOPIC: 11.1, 11.3, 11.4, 11.5 (11.5.6).** Rate of Reaction (308-313), Energy of Activation, Finding of Order of Reaction. (Pg. 316-319). Arrhenius Equation (Pg. 322-324).  
 Classwork: Q.1, 2, Q.3 (i, ii, iv, v), Q.8, Q.19, Q.20, Q.21, Q.22.  
 Homework: Q.4, Q.5, Q.6, Q.7 (i, iii, iv), Q.9, Q.15.  
**LIST OF EXPERIMENTS (CHEMISTRY) PART-I**

- Crystallization of benzoic acid from water.
- To separate a mixture of various inks by paper chromatography.
- Separation and Identification of lead and cadmium ions in a mixture solution by paper chromatography.
- Determination of heat of neutralization of NaOH and HCl.
- Preparation of standard solution of alkalis and acids e.g., NaOH, KOH, Oxalic acid, succinic acids.
- Preparation of solution of H<sub>2</sub>SO<sub>4</sub> of approximate strength and then determination of its exact strength with the help of standard Na<sub>2</sub>CO<sub>3</sub> solution.
- To prepare a standard solution of oxalic acid and standardize a solution of NaOH.
- To determine the solubility of oxalic acid at room temperature. You are provided with 0.1 M NaOH.
- Determination of acetic acid in vinegar.
- The given solution contains 15 g of mixture of NaOH and Na<sub>2</sub>SO<sub>4</sub> per dm<sup>3</sup>. Calculate the amount of NaOH in 45 grams of the mixture. 0.1 M HCl is given.
- Determination of free alkali in soap.
- Determination of Na<sub>2</sub>CO<sub>3</sub> in washing soda.
- Determination of percentage of purity of Na<sub>2</sub>CO<sub>3</sub> in the given solution containing 10 g. of impure Na<sub>2</sub>CO<sub>3</sub> sample/dm<sup>3</sup>. You are provided with 0.1 M HCl solution.
- 28.6 grams of washing soda (Na<sub>2</sub>CO<sub>3</sub> · xH<sub>2</sub>O) have been dissolved/dm<sup>3</sup>. Calculate the number of water molecules of crystallization. You are provided with 0.1 M HCl solution.
- Determination of NaHCO<sub>3</sub> in the given sample of baking soda. 0.1M HCl soln. is provided.
- 8.4 gram M HCO<sub>3</sub> are dissolved per dm<sup>3</sup> of solution. Find out At. Wt. of M. 0.05 M H<sub>2</sub>SO<sub>4</sub> is given.
- You are given the solution of KMnO<sub>4</sub>. Calculate its volume required to prepare 1.0 dm<sup>3</sup> of 0.002M KMnO<sub>4</sub> solution.
- The given soln. 'A' contains 10 grams of a mixture of H<sub>2</sub>SO<sub>4</sub> and oxalic acid dissolved/dm<sup>3</sup>. Determine the percentage of H<sub>2</sub>SO<sub>4</sub> in the mixture. 0.02M KMnO<sub>4</sub> is given.
- Determine the no. of molecules of water of crystallization in a given sample of oxalic acid by permanganate titration. The amount of oxalic acid dissolved per dm<sup>3</sup> is 6.3 g.
- Determination of solubility of oxalic acid at room temperature.
- To determine the strength of ferrous sulphate solution by titrating it against 0.02M KMnO<sub>4</sub>.
- The given solution contains 30 gram of partially oxidized FeSO<sub>4</sub> · 7H<sub>2</sub>O dissolved per dm<sup>3</sup>. Determine the %age of oxidation of the given sample.
- To determine the strength of given ferrous ammonium sulphate (Mohr's salt) by titrating it against standard potassium permanganate solution.
- The given solution contains 40g. of FeSO<sub>4</sub>(NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub> · xH<sub>2</sub>O. dissolved per dm<sup>3</sup>. Determine the value of x.
- Determine the solubility of given sample of Mohr's salt at room temperature. You are provided with 0.02M KMnO<sub>4</sub>.
- Prepare a standard (M/10) 250 cm<sup>3</sup>. Solution of iodine. 0.1 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> is provided.
- 24.8 grams of a sample of alkali thiosulphate (M<sub>2</sub>S<sub>2</sub>O<sub>3</sub>) are dissolved in 1 dm<sup>3</sup> of the given solution. Calculate the atomic weight of the metal by a volumetric method. Given M/10 iodine solution.
- 20 gram of Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> are dissolved in one dm<sup>3</sup> solution. Find out the %age of sulphur. You are provided with 0.05M iodine solution.

## CHAPTER-1 (BASIC CONCEPTS)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS-2020)

- The atomic masses of elements were determined by  
 (a) John Dalton (b) J. Berzelius (c) Soddy (d) Moseley  
 (16 times)
- Mass of one mole of electrons is:  
 (a) 0.55 mg (b) 0.184 mg (c) 1.673 mg (d) 1.008 mg
- The mass of two moles of electrons is:  
 (a) 1.10 mg (b) 1.008 mg (c) 0.184 mg (d) 1.673 mg
- Tin has isotopes:  
 (a) 7 (b) 9 (c) 11 (d) 5
- Isotopes differ in number of:  
 (a) Electron (b) Neutron (c) Proton (d) Proton Number
- How many isotopes are present in palladium.  
 (a) four (b) Five (c) Six (d) Seven
- The element nickel has isotopes  
 (a) 5 (b) 2 (c) 7 (d) 3
- Bromine has isotopes:  
 (a) 8 (b) 6 (c) 4 (d) 2
- Cadmium has isotopes:  
 (a) 3 (b) 4 (c) 5 (d) 9
- The number of neutron present in <sup>39</sup><sub>19</sub>K is:  
 (a) 39 (b) 18 (c) 20 (d) 19
- The number of moles of CO<sub>2</sub> which contains 8.0g of oxygen is: (19 times)  
 (a) 0.25 (b) 0.50 (c) 1.0 (d) 1.50
- The number of moles of CO<sub>2</sub> which contains 16g of Oxygen is: (2 times)  
 (a) 0.25 (b) 0.50 (c) 1.0 (d) 1.50
- One mole of SO<sub>2</sub> contain (12 times)  
 (a) 6.02 × 10<sup>23</sup> atoms of Oxygen (b) 18.01 × 10<sup>23</sup> molecules of SO<sub>2</sub>  
 (c) 6.02 × 10<sup>23</sup> atoms of Sulphur (d) 4 gram atoms of SO<sub>2</sub>
- The mass of CO<sub>2</sub> containing 8 grams of oxygen (O<sub>2</sub>) in gram is:  
 (a) 32 (b) 22 (c) 16 (d) 11
- Amount of NaOH required to produce 250cm<sup>3</sup> of 1M solution in grams is:  
 (a) 10 (b) 15 (c) 20 (d) 25
- The largest number of molecules are present in:  
 (a) 2.8g of CO (b) 3.6g of H<sub>2</sub>O (c) 4.6g of C<sub>2</sub>H<sub>5</sub>OH (d) 5.4g of N<sub>2</sub>O<sub>5</sub>
- The number of atoms present in 0.1 mole of oxygen gas are:  
 (a) 6.02 × 10<sup>22</sup> (b) 3.01 × 10<sup>23</sup> (c) 2 × 6.02 × 10<sup>22</sup> (d) 9.03 × 10<sup>22</sup>
- The volume occupied by 1.4g of N<sub>2</sub> at S.T.P in dm<sup>3</sup> is: (14 times)  
 (a) 1.12 dm<sup>3</sup> (b) 2.24 dm<sup>3</sup> (c) 11.2 dm<sup>3</sup> (d) 22.4 dm<sup>3</sup>
- The volume occupied by 32g of O<sub>2</sub> at S.T.P is:  
 (a) 22.414 dm<sup>3</sup> (b) 2.2414 dm<sup>3</sup> (c) 224.414 dm<sup>3</sup> (d) 0.224 dm<sup>3</sup>
- Percentage of nitrogen in NH<sub>3</sub> is: (1 time)  
 (a)  $\frac{14}{17} \times 100$  (b)  $\frac{14}{34} \times 100$  (c)  $\frac{3}{17} \times 100$  (d)  $\frac{3}{34} \times 100$
- Fractional atomic mass is mainly due to:  
 (a) Mass of atom is in fraction (b) Atomic mass is average mass of isobars  
 (c) Elements mostly consist of isotopes having different fractional abundances  
 (d) Atomic mass is average masses of isotopes
- The volume occupied by 16 g of CH<sub>4</sub> at STP is:  
 (a) 2.24 dm<sup>3</sup> (b) 22.414 dm<sup>3</sup> (c) 1.3 dm<sup>3</sup> (d) 1.8 dm<sup>3</sup>
- Volume occupied by one mole of gas at standard temperature and pressure is:  
 (a) 54 dm<sup>3</sup> (b) 22.414 dm<sup>3</sup> (c) 2.24 dm<sup>3</sup> (d) 2.4 dm<sup>3</sup>



## Answers

1	2	3	4	5	6	7	8	9	10	11	12	13
b	a	a	c	b	c	a	d	d	c	a	b	c
14	15	16	17	18	19	20	21	22	23			
d	a	b	c	a	a	a	d	b	b			

## CHAPTER-1 (BASIC CONCEPTS)

### SHORT QUESTION'S

(ACCORDING TO ALP SMART SYLLABUS-2020)

1. Isotopes:

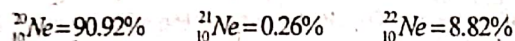
1. What are isotopes? Why they have same chemical but different physical properties? (2 times)

Ans: Atoms of the same elements can possess different atomic masses but same atomic numbers, such atoms of an element are called isotopes. For example carbon has three isotopes written as  $^{12}_6\text{C}$ ,  $^{13}_6\text{C}$  and  $^{14}_6\text{C}$  expressed as C-12, C-13 and C-14.

Isotopes of an elements have same chemical properties and same position in the periodic table, but due to different atomic masses have different physical properties.

2. Explain relative abundance of isotopes with suitable example? (1 time)

Ans: The percentage of isotope of an element in comparison to other isotopes of the same element is called its relative abundance. It is measured by mass spectrometry. e.g. Neon



3. How do no individual Ne atom in the sample of the element has mass of 20.18 a.m.u?

Ans: The overall atomic mass of neon, which is an ordinary isotopic mixture, is the average of the determined atomic masses of individual isotopes. Hence

$$\text{Average atomic mass} = \frac{20 \times 90.92 + 21 \times 0.26 + 22 \times 8.82}{100} = 20.18 \quad \text{Answer}$$

Hence the average atomic mass of neon is 20.18 amu.

It is important to realize that no individual neon atom in the sample has a mass of 20.18 amu.

4. How the isotopes of an element are separated by mass spectrometer?

Ans: Isotopes of an element are separated on the basis of their m/e ratio.

5. Define Isotope. Give an example?

Ans: Atoms of the same elements can possess different atomic masses but same atomic numbers, such atoms of an element are called isotopes. For example carbon has three isotopes written as  $^{12}_6\text{C}$ ,  $^{13}_6\text{C}$  and  $^{14}_6\text{C}$  expressed as C-12, C-13 and C-14.

6. Write only names of any four methods employed for the separation of Isotopes.

Ans: The separation of isotopes can be done by the methods based on their properties. Some important methods are, gaseous diffusion, thermal diffusion, distillation, ultracentrifuge, electromagnetic separation and laser separation.

7. 4.9g of  $\text{H}_2\text{SO}_4$  when completely ionized in water have equal number of positive and negative charges but the number of positively charged ions are twice the number of negatively charged ions? (2 times 2018)

Ans:  $\text{H}_2\text{SO}_4 \rightleftharpoons 2\text{H}^+ + \text{SO}_4^{2-}$

This equation shows that 1 mole of  $\text{H}_2\text{SO}_4$  produce 2 moles of positive ion and 1 mole of negative ions. But the number of charges are same. Hence whatever the amount of  $\text{H}_2\text{SO}_4$  is the number of positive ions will be twice the number of negative ions.

5. Mole and Avogadros Number:

8. 180g of glucose and 342g of sucrose have the same number of molecules but different number of atoms present in them. Explain? (5 times)

Ans: Mass of glucose = 180 g

$$\text{Number of moles of glucose} = \frac{\text{Mass of glucose}}{\text{Molar mass of glucose}} = \frac{180}{180} = 1 \text{ mole}$$

Mass of sucrose = 342 g

$$\text{Number of moles of sucrose} = \frac{\text{Mass of sucrose}}{\text{Molar mass of sucrose}} = \frac{342}{342} = 1 \text{ mole}$$

And one mole of each compound contains  $6.02 \times 10^{23}$  numbers of particles. So both have same number of particles. Glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) molecule have 24 atoms, while sucrose ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) molecule have 45 atoms. So equal number of molecules of glucose and sucrose will have different number of atoms.

9. Calculate the number of  $\text{CO}_2$  molecules in 20.0 grams of it?

Ans: No. of  $\text{CO}_2$  molecules (N) = ?

Given mass (m) = 20g

Molar mass of  $\text{CO}_2$  = 44g/mole

Avogadro's number ( $N_A$ ) =  $6.022 \times 10^{23}$

$$\text{Formula } N = \frac{m}{M} \cdot N_A = \frac{20}{44} \times 6.02 \times 10^{23} = 2.73 \times 10^{23}$$

No. of  $\text{CO}_2$  molecules are  $2.73 \times 10^{23}$

10. Mg atom is twice heavier than that of carbon atom. Comment?

Ans: As (8 times) (Class work)

12g of C = 1 mole of C =  $6.02 \times 10^{23}$  atoms of carbon

24 g of Mg = 1 mole of Mg =  $6.02 \times 10^{23}$  atoms of magnesium

Equal number of magnesium atoms have twice mass as compared to equal number of carbon atoms. Therefore each magnesium atom is twice heavier than carbon atom.

11. One Mole of water has two moles of bonds, three moles of atoms. Explain? (2 times) (Class work)

Ans: In one mole of water, moles of oxygen = 1 mole

In one mole of water, moles of hydrogen = 2 mole

In one mole of water, number of bonds = 2 mole

In one mole of water, number of electrons (8 electron of O and two electron of H) = 10 moles

In one mole of water, number of particle of oxygen ( $8P+8N+8e^-$ ) = 24 moles

In one mole of water, number of particle hydrogen =  $2P+2e^-$  = 4 moles

So, total number of particles in one mole of water =  $24+4$  = 28 moles

1 mol of  $\text{H}_2\text{O}$  = 2 mol H atom + 1 mole of oxygen atom = 3 moles

12. Define mole and molar volume? (4 times)

Ans: The atomic mass of an element or molecular mass of a molecule or formula mass of a formula unit is expressed in grams is called gram atomic mass, gram molecular mass and gram formula unit respectively. These are also called gram mole or simply mole.

One mole of a gas at standard temperature and pressure (STP) occupies a volume of  $22.414 \text{ dm}^3$  is called molar volume and it is true only when the gas is ideal.

13. What do you know about gram atom? (2 times)

Ans: When the substance at our disposal is an element then the atomic mass of that element expressed in grams is called one gram atom. It is also called one gram mole or simply a mole of that element.

$$\text{Number of gram atoms or moles of an element} = \frac{\text{Mass of an element in grams}}{\text{Molar mass of the substance}}$$

For example

$$\begin{aligned} 1 \text{ gram atom of hydrogen} &= 1.008 \text{ g} \\ 1 \text{ gram atom of carbon} &= 12.000 \text{ g} \\ 1 \text{ gram atom of uranium} &= 238.0 \text{ g} \end{aligned}$$



14. How many oxygen atoms are present in 4.8g of Ozone (O=16 amu)?

Ans: Molecular mass of  $O_3 = 16 \times 3 = 48 \text{ gmol}^{-1}$   
 $\text{Mass in grams} = 4.8 \text{ g}$   
 $\text{Moles of } O_3 = \frac{\text{Mass in grams}}{\text{Formula mass}} = \frac{4.8}{48} = 0.1 \text{ mol}$

1 mole of  $O_3$  has 3 moles of oxygen atoms.  
 0.1 moles of  $O_3$  has  $= 0.1 \times 3 = 0.3 \text{ moles}$

15. 23g of Na and 238g of U have equal number of atoms in them. Justify? (2 time) (Class work)

Ans: Mass of Na = 23 g

Number of moles of Na  $= \frac{\text{Mass of Na}}{\text{Molar mass of Na}} = \frac{23}{23} = 1 \text{ mole}$

Mass of U = 238 g

Number of moles of U  $= \frac{\text{Mass of U}}{\text{Molar mass of U}} = \frac{238}{238} = 1 \text{ mole}$

Since one mole of each element contains  $6.02 \times 10^{23}$  particles.

16. Calculate mass in Kg of  $2.6 \times 10^{20}$  molecules of  $SO_2$ ?

Ans: No. of molecules of  $SO_2(N) = 2.6 \times 10^{20}$   
 Molar mass of  $SO_2(M) = 32 + 32 = 64 \text{ g/mole}$   
 Mass in Kg of  $SO_2(m) = ?$

$$N = \frac{m}{M} \cdot N_A$$

$$m = \frac{N \times M}{N_A} = \frac{2.6 \times 10^{20} \times 64}{6.02 \times 10^{23}}$$

$$m = 27.64 \times 10^{-3} \text{ g} = 27.64 \times 10^{-6} \text{ Kg}$$

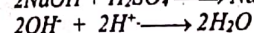
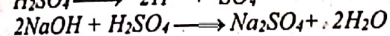
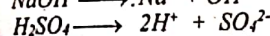
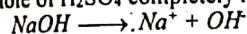
17. Calculate the number of moles of O atoms in 9.00 gram of  $Mg(NO_3)_2$ ?

Ans: Formula mass of  $Mg(NO_3)_2 = 24 + 14 \times 2 + 16 \times 6 = 148 \text{ gmol}^{-1}$   
 $\text{Mass in grams} = 9 \text{ g}$   
 $\text{Moles of } Mg(NO_3)_2 = \frac{\text{Mass in grams}}{\text{Formula mass}} = \frac{9}{148} = 0.06 \text{ mol}$

1 mole of  $Mg(NO_3)_2$  has 6 moles of oxygen atoms.  
 0.06 moles of  $Mg(NO_3)_2$  has  $= 0.06 \times 6 = 0.36 \text{ moles}$

18. One mole of  $H_2SO_4$  should completely react with two moles of NaOH. How does Avogadro's number help to explain it? (2 times)

Ans: One mole of  $H_2SO_4$  completely neutralize with two moles of NaOH



1 mole of  $H_2SO_4$  when completely ionizes produces two moles of  $H^+$  ions or  $2 \times 6.02 \times 10^{23}$  ions. Two mole of NaOH also produces two moles of  $OH^-$  ions or  $2 \times 6.02 \times 10^{23}$  ions. When these solution are mixed together, they have equal number of  $H^+$  and  $OH^-$  ions and solution becomes neutral.

19. Calculate the moles of 100g of silicon? (Atomic Mass of Si is 28) (2 times)

Ans: Given mass of silicon (m) = 100g  
 Molar mass of silicon (M) = 28g/mole  
 No. of moles (n) = ?

$$n = \frac{m}{M} = \frac{100}{28} = 3.56 \text{ moles}$$

20. Calculate the moles of chlorine atoms in 0.822g  $C_2H_4Cl_2$ ? (1 time)

Ans: Given mass (m) = 0.822g  
 Number of mole of Cl atom (n) = ?  
 Molar mass of  $C_2H_4Cl_2$  (M) = 99 g  $\text{mol}^{-1}$

$$\text{No. of moles of } C_2H_4Cl_2 = \frac{0.822}{99} = 0.00830 = 8.30 \times 10^{-3}$$

One mole of  $C_2H_4Cl_2 = 2 \text{ moles of Cl atoms}$

$8.30 \times 10^{-3} \text{ moles of } C_2H_4Cl_2 = 2 \times 8.30 \times 10^{-3} = 0.017 \text{ moles of Cl atoms.}$

21. Calculate number of gram atoms of Na when its mass is 0.1kg? (2 times)

Atomic mass of Na = 23g  $\text{mol}^{-1}$

Ans: Gram atoms of Na = ?

Given mass of Na = 0.1 kg = 100 g

Atomic mass of Na = 23 g/mole

$$\text{Formula } n = \frac{m}{M} = \frac{100 \text{ g}}{23 \text{ gmol}^{-1}} = 4.34 \text{ moles}$$

The gram atoms(moles) of Na is 4.34 moles.

22. How many molecules of  $H_2O$  are present when its amount is 0.25 moles?

Ans: No. of  $H_2O$  molecules (N) = ?

No. of moles of  $H_2O(n) = 0.25 \text{ mole}$

Avogadro's ( $N_A$ ) =  $6.02 \times 10^{23}$

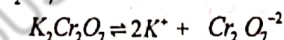
Number of water molecule =  $n \times N_A$

$$= 0.25 \times 6.02 \times 10^{23}$$

$$\text{No. of } H_2O \text{ molecules} = 1.505 \times 10^{23} \text{ molecules}$$

23. One mole of  $K_2Cr_2O_7$  has thrice the number of ions than the number of formula units when ionized? (Class work)

Ans:  $K_2Cr_2O_7$  ionize as:



This equation show that 1 mole of  $K_2Cr_2O_7$  produces 2 moles of  $K^+$  and 1 mole of  $Cr_2O_7^{2-}$ .

Hence total three moles of ions are produced by the ionization of 1 mole of  $K_2Cr_2O_7$ .

24. 100cm<sup>3</sup> of  $NH_3$  gas and 100cm<sup>3</sup> of  $H_2$  gas at STP contain equal number of molecules. Justify it?

Ans: Different gases have different molecular masses. According to Avogadro's law "Equal volumes of different gases contain equal number of molecules." As there are large spaces between gas molecule so molecular gases and molecular sizes do not effect the overall occupied volume.

25. Calculate the mass in grams of 2.74 moles of  $KMnO_4$ ? Formula mass of  $KMnO_4$  is 158 g  $\text{mol}^{-1}$  (2 time)

Ans: Moles of  $KMnO_4(n) = 2.74 \text{ mole}$

Molar mass (M) =  $39 + 55 + 16 \times 4 = 158 \text{ g/mole}$

1 mole of  $KMnO_4 = 158 \text{ g}$

2.74 Moles of  $KMnO_4 = 2.74 \times 158 = 432.92 \text{ g}$

26. Calculate number of molecules in 9 g of ice? (3 times)

Ans: Given mass of  $H_2O(m) = 9 \text{ g}$

Molar mass of  $H_2O(M) = 18 \text{ g/mole}$

No. of molecules of  $H_2O$  in 9g of ice (n) = ?

$$N = \frac{m}{M} \cdot N_A = \frac{9}{18} \times 6.02 \times 10^{23} = 3.01 \times 10^{23} \text{ molecules}$$

27. How has one mg of  $K_2Cr_2O_7$  thrice the number of ions than the number of formula units when ionized in water? (Class work)

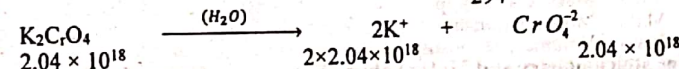
Ans: it can be justified as following

Mass of  $K_2Cr_2O_7 = 1 \text{ mg} = 0.001 \text{ g}$

Molar mass of  $K_2Cr_2O_7 = 294 \text{ g}$

$$\text{Number of formula units of } K_2Cr_2O_7 = \frac{\text{Mass} \times N_A}{\text{Molar mass}}$$

$$= \frac{0.001 \times 6.02 \times 10^{23}}{294} = 2.04 \times 10^{18}$$



$$2.04 \times 10^{18}$$

$$2 \times 2.04 \times 10^{18}$$

$$2.04 \times 10^{18}$$



- Hence it is justified that total number of ions is thrice than number of formula units ionized.
28. Why do 2g of  $H_2$ , 16g of  $CH_4$ , 44g of  $CO_2$  occupy separately the volume of 22.414 dm<sup>3</sup> although the sizes and masses of molecules of three gases are very different from each other? (Class work)
- Ans: 2g of  $H_2$  = 1 mole =  $6.02 \times 10^{23}$  molecules = 22.414 dm<sup>3</sup> volume at STP.  
 16g of  $CH_4$  = 1 mole =  $6.02 \times 10^{23}$  molecules = 22.414 dm<sup>3</sup> volume at STP.  
 44g of  $CO_2$  = 1 mole =  $6.02 \times 10^{23}$  molecules = 22.414 dm<sup>3</sup> volume at STP.  
 According to Avogadro's law, equal numbers of molecules of all gases occupy same volumes at same temperature and pressure. Since,  $H_2$ ,  $CH_4$  and  $CO_2$  have same number of molecules that is why these occupy same volume.
29. Calculate mass of  $10^{-3}$  moles of  $MgSO_4$ .
- Ans:  $MgSO_4$  is an ionic compound. We will consider its formula mass in place of molecular mass.
- Number of gram formula or moles of a substance =  $\frac{\text{Mass of the ionic substance}}{\text{Formula mass of the ionic substance}}$
- Formula mass of  $MgSO_4$  = 24+96=120 gmol<sup>-1</sup>  
 Number of moles of  $MgSO_4$  =  $10^{-3}$  moles  
 Applying the formula
- $$10^{-3} = \frac{\text{Mass of } MgSO_4}{120 \text{ gmol}^{-1}}$$
- Mass of  $MgSO_4$  =  $10^{-3} \text{ moles} \times 120 \text{ gmol}^{-1}$   
 =  $120 \times 10^{-3} = 0.12 \text{ g}$  Answer
30. Define Avogadro's number. Give its numerical value? (4 times) (Home work)
- Ans: Avogadro's number: The number of particles (atoms, ions or molecules) present in mole of a substance is called Avogadro's number. Its value is  $6.02 \times 10^{23}$
6. **Stoichiometry and Limiting Reactants:**
31. Write down limitations of a chemical equation.
- Ans: Chemical equations have certain limitations. They do not tell about the conditions and the rate of reaction. Chemical equation can even be written to describe a chemical change that does not occur.
32. Law of conservation of mass has to be obeyed during stoichiometric calculations. Justify it. (3 times) (Class work)
- Ans: Stoichiometric calculations depend upon balanced chemical equation, while the balanced chemical equation is written according to law of conservation of mass. So therefore law of conservation of mass is to be obeyed during stoichiometric calculations.
33. List steps involved to identify a limiting reactant? (5 times)
- Ans: To identify a limiting reactant the following steps are performed:  
 1) Calculate the number of moles from given amount of reactant.  
 2) Find out the number of moles of product with the help of a balanced chemical equation.  
 3) Identify the reactant which produces the least amount of product as limiting reactant.
34. Write two assumptions of stoichiometry? (4 times)
- Ans: There are following assumptions of stoichiometry:  
 1) All the reactants are completely converted into products.  
 2) No side reaction occurs.  
 3) While doing calculations, the law of conservation of mass and law of definite proportions are obeyed.
35. By using a balanced chemical equation. What type of relationships can be studied?
- Ans: The following types of relationships can be studied with the help of balanced chemical equation.  
 i) Mass - Mass relationship  
 ii) Mass - Mole relationship  
 iii) Mass - Volume relationship
36. Define stoichiometry and Molar volume? (4 times) (Home work)

- Ans: Stoichiometry: A branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.
38. 11g of carbon is reacted with 32g of  $O_2$  to give  $CO_2$ . Which one is the limiting reactant?
- Ans:  $C + O_2 \rightarrow CO_2$
- According to balanced chemical equation
- $$C = 1 \text{ mole}$$
- $$O_2 = 1 \text{ mole}$$
- $$CO_2 = 1 \text{ mole}$$
- But we are provided with moles of C =  $\frac{m}{M} = \frac{11}{12} = 0.91 \text{ g mole}^{-1}$
- Moles of  $O_2 = \frac{m}{M} = \frac{32}{32} = 1 \text{ mole}$
- Hence "C" has less number of moles. So, it is a limiting reactant.
39. Define limiting reactant, give an example? (6 times)
- Ans: The limiting reactant is a reactant that control the amount of product formed in a chemical reaction due to its smaller amount. e.g.  $2H_2 + 2O_2 \rightarrow 2H_2O$ . In this reaction  $H_2$  is a limiting reactant.
40. How do many chemical reactions take place in our surroundings involve the limiting reactant? (Class work)
- Ans: Many chemical reactions taking place in our surrounding involve the limiting reactants.
- Example:
- Combustion is a daily life reaction in which fuel is the limiting reactant and oxygen is non limiting reactant. Once the fuel like  $CH_4$ , will consume the combustion will stop.
  - Rusting of iron is another example. In this reaction iron is the limiting reactant, once it will completely convert in to rust, the further reaction will stop.
41. Many chemical reactions taking place in our surroundings involve the limiting reactants. Explain with example? (3 times) (Class work)
- Ans: Example: Burning of coal
- Explanation:
- In burning of coal oxygen ( $O_2$ ) and coal (C) are reactants. As oxygen is present in excess amount so coal (C) is our limiting reactants.
- $$C + O_2 \rightarrow CO_2$$
- (Excess)
7. **Yield:**
42. What is percentage yield? Give its significance? (2 times) (Home work)
- Ans: A chemist is usually interested in the efficiency of a reaction. The efficiency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage yield.
- $$\% \text{ yield} = \frac{\text{Actual yield}}{\text{theoretical yield}} \times 100$$
43. Differentiate between actual (Experimental) yield and theoretical yield? (2 times)
- Ans: The amount of the products obtained in chemical reaction is called the actual yield of that reaction.  
 The amount of the products calculated from the balanced chemical equation represents the theoretical yield. The theoretical yield is the maximum amount of the product that can be produced by a given amount of a reactant, according to the balanced chemical equation.
44. Actual yield is usually less than theoretical yield. Why? (13 times)
- Ans: Actual yield of a chemical reaction is always less than its theoretical yield due to following reasons:



- Loss of product due to inexperienced work, the product is lost in separation techniques e.g. crystallization, distillation, filtration etc.  
 i) Some of reactants might take part in competing side reactions.  
 ii) Due to reversibility of reactions.  
 iii) How can the efficiency of a chemical reaction be expressed? (3 times)  
 The efficiency of a reaction is expressed in terms of %age yield.

Ans: 
$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

46. Calculate %age of Nitrogen in  $\text{NH}_2\text{CO}\text{NH}_2$  (at mass of N=14, C=12 H=1, O=16) (2 times)  
 Ans: Molecular mass of  $\text{NH}_2\text{CO}\text{NH}_2 = 14 + 2(1) + 12 + 16 + 14 + 2(1) = 60$  amu  
 %age of Nitrogen in  $\text{NH}_2\text{CO}\text{NH}_2 = \frac{28}{60} \times 100 = 46.66\%$

2019

47. Define Relative atomic mass. Give two examples.  
 Ans: Relative atomic mass is mass of an atom of an element as compared to mass of an atom of carbon taken as 12. The unit of relative atomic mass is a.m.u.  
 For example: Relative atomic mass of hydrogen is 1.008 a.m.u.  
 Relative atomic mass of oxygen = 15.99 amu
48. Define Gram formula with one example.  
 Ans: "Formula unit mass of an ionic compound expressed in grams is called one gram-formula or one mole of that ionic compound."  
 For example: 1 gram-formula (1mole) of  $\text{NaCl} = 58.5 \text{ gmol}^{-1}$
49. How many atoms are present in 0.1 g of Na - 23?

Ans: 
$$\text{No. of atoms} = \frac{\text{Mass of element in g}}{\text{Molar mass}} \times N_A$$

$$= \frac{0.1 \text{ g}}{23 \text{ gmol}^{-1}} \times 6.02 \times 10^{23}$$

$$= 2.6 \times 10^{21} \text{ atoms}$$

50. Why mass of Ne gas is in fractions.  
 Ans: Atomic mass Neon is 20.18 amu. Actually, atomic mass of an element is taken as average of atomic masses of all of its isotopes proportional to their relative abundances. Neon has three isotopes, Ne-20, Ne-21, Ne-22 in %ages 90.92%, 0.26%, 8.82% respectively.
51.  $\text{N}_2$  and CO have same no. of protons and Neutrons. Justify. (Class work)

Ans:

	Protons	Neutrons
$^{14}_7\text{N}$	7	7
$\text{N}_2$	14	14
$^{12}_6\text{C}$	6	6
$^{16}_8\text{O}$	8	8
CO	14	14

Hence, it is justified that both  $\text{N}_2$  and CO have same no. of protons and Neutrons.

52. What is Justification of two strong peaks of almost equal heights in mass spectrum of Bromine.

Ans: Two strong peaks of almost equal length in mass spectrum of Bromine means, it has two isotopes of nearly equal %age abundance. e.g; Br - 80 with 50.54% abundance and Br - 81 with 49.49% abundance.

53. What are monoisotopic elements?

Ans: Element having single isotope is called monoisotopic element.  
 For example: Iodine (I), Gold (Au) etc.

## CHAPTER-1 (BASIC CONCEPTS) LONG QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

- Explain isotopes with their relative abundance.
- Define the following and give one example of each.  
 (i) Mole (ii) Isotopes (iii) Molecular ion (iv) Avogadro's number
- Serotonin (Molar Mass=176g/mol) is a compound that conducts nerve impulses in brain and muscles. It contains 68.2%C, 6.86%H, 15.9%N and 9.08%O. What is its Molecular Formula? (3 times)
- Write a note on Avogadro's number. (2 times)
- A well known ideal gas is enclosed in a container having volume  $500 \text{ cm}^3$  at STP its mass comes out to be 0.72g. What is the molar mass of this gas. (3 times)
- Calculate the number of grams of  $\text{Al}_2\text{S}_3$  which can be prepared by the reaction of 20g of Al & 30g of sulphur. How much non-limiting reactant is in excess? (1 time)
- A mixture of  $\text{NH}_3$  gas can be prepared by heating together two solids  $\text{NH}_4\text{Cl}$  and  $\text{Ca(OH)}_2$ . If a mixture containing 100g of each solid is heated then calculate the number of grams of  $\text{NH}_3$  produced.  
 $4\text{NH}_4\text{Cl}_{(s)} + \text{Ca(OH)}_{2(s)} \rightarrow \text{CaCl}_{2(s)} + 2\text{NH}_{3(g)} + 2\text{H}_2\text{O}_{(l)}$  (1 time)(2017=1 time)  
 Molar mass of  $\text{NH}_4\text{Cl} = 53.5 \text{ gmol}^{-1}$  and  $\text{Ca(OH)}_2 = 74 \text{ gmol}^{-1}$  (1 time)(2016=1)
- A mixture of two liquids hydrazine  $\text{N}_2\text{H}_4$  and  $\text{N}_2\text{O}_4$  are used in rockets. They produce  $\text{N}_2$  and water vapours. How many grams of  $\text{N}_2$  gas will be formed by reacting 100 g of  $\text{N}_2\text{H}_4$  and 200 g of  $\text{N}_2\text{O}_4$ . ( $2\text{N}_2\text{H}_4 + \text{N}_2\text{O}_4 \rightarrow 4\text{H}_2\text{O} + 3\text{N}_2$ )
- Calculate the number of grams of  $\text{K}_2\text{SO}_4$  and water produced when 14 g of KOH are reacted with excess of  $\text{H}_2\text{SO}_4$ ?
- Mg metal reacts with HCl to give hydrogen gas. What is the minimum volume of HCl solution (27% by weight required to produce 12.1 g  $\text{H}_2$ ). The density of HCl solution is  $1.14 \text{ g cm}^{-3}$ .  
 $4\text{Mg}_{(s)} + 2\text{HCl}_{(aq)} \rightarrow \text{MgCl}_{2(aq)} + \text{H}_{2(g)}$   
 At. Mass of Mg = 24g / mol At. Mass of Cl = 35.5 g / mol (3 times)(2017=1 time)
- Calculate the number of grams of  $\text{K}_2\text{SO}_4$  and water produced when 14g of KOH are reacted with excess of  $\text{H}_2\text{SO}_4$ . Also calculate the number of molecules of water produced:  
 $2\text{KOH}_{(aq)} + \text{H}_2\text{SO}_{4(aq)} \rightarrow \text{K}_2\text{SO}_{4(aq)} + 2\text{H}_2\text{O}_{(l)}$   
 (At. Mass of K = 39, S = 32, O = 16) (2 times)
- Define Stoichiometry. Give assumptions, Mention any two important laws which help to perform stoichiometric calculation.
- What is difference between actual yield and theoretical yield? Why actual yield is less than theoretical yield? (3 times)
- When lime stone ( $\text{CaCO}_3$ ) is roasted quick lime (CaO) is Produced according to the equation.  
 The actual yield of CaO is 2.5 Kg when 4.5 Kg quick lime of lime stone is roasted. What is the percentage yield of this reaction  
 $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$  Atomic mass of Ca = 40, C = 12, O = 16 (4 times)
- Silicon Carbide ( $\text{SiC}$ ) is an important ceramic material. It is produced by allowing sand ( $\text{SiO}_2$ ) to react with Carbon at high temperature.  $\text{SiO}_2 + 3\text{C} \rightarrow \text{SiC} + 2\text{CO}$   
 When 100 kg sand is reacted with excess of Carbon, 51.4 kg of SiC is produced. (2 times)
- What is the percentage yield of SiC?
- Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7% hydrogen and 51.6% oxygen. Its molar mass is  $62.1 \text{ gmol}^{-1}$ . Determine its empirical and molecular formula. (4 times)



# Solved Exercise Chapter # 1

## ALP SMART SYLLABUS 2020

Q1 Select the most suitable answer from the given ones in each question. (Class work)

- (i) Isotopes differ in  
 (a) properties which depend upon mass  
 (b) arrangement of electrons in orbitals  
 (c) chemical properties  
 (d) the extent to which they may be affected in electromagnetic field.
- (ii) Select the most suitable answer from the given ones in each question.  
 (a) Isotopes with even atomic masses are comparatively abundant.  
 (b) Isotopes with odd atomic masses are comparatively less abundant.  
 (c) Isotopes with even atomic masses and even atomic numbers are comparatively abundant.  
 (d) Isotopes with even atomic masses and odd atomic numbers are comparatively abundant.
- (iii) Many elements have fractional atomic masses. This is because  
 (a) the mass of the atom is itself fractional.  
 (b) atomic masses are average masses of isobars.  
 (c) atomic masses are average masses of isotopes.  
 (d) atomic masses are average masses of isotopes proportional to their relative abundance.
- (iv) 27 g of Al will react completely with how much mass of O<sub>2</sub> to produce Al<sub>2</sub>O<sub>3</sub>.  
 (a) 8 g of oxygen (b) 16 g of oxygen (c) 32 g of oxygen (d) 24 g of oxygen
- (v) A limiting reactant is the one which  
 (a) is taken in lesser quantity in grams as compared to other reactants.  
 (b) is taken in lesser quantity in volume as compared to the other reactants.  
 (c) gives the maximum amount of the product which is required.  
 (d) gives the minimum amount of the product under consideration.

Q2. Fill in the blanks. (Class work)

- (i) The unit of relative atomic mass is a.m.u.  
 (ii) The exact masses of isotopes can be determined by mass spectrograph.  
 (iii) The phenomenon of isotopy was first discovered by Soddy.  
 (iv) A limiting reagent is that which controls the quantities of product.  
 (v) 4g of CH<sub>4</sub> at 0°C and 1 atm pressure has  $1.505 \times 10^{23}$  molecules of CH<sub>4</sub>.  
 (vi) Stoichiometric calculations can be performed only when law of conservation of mass is obeyed.

Q3. Indicate true or false as the case may be: (Class work)

- (i) Neon has three isotopes and the fourth one with atomic mass 20.18 amu. (False)  
 (ii) The number of atoms in 1.79 g of gold and 0.023 g of sodium are equal. (False)  
 (iii) The number of electrons in the molecules of CO and N<sub>2</sub> are 14 each, so 1 g of each gas will have same number of electrons. (True)  
 (iv) Actual yield of a chemical reaction may be greater than the theoretical yield. (False)

Q6 Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e. Ag-107 and Ag-109. Given the following mass spectrometric data, calculate the average atomic mass of silver. Isotopes Mass (amu) Percentage abundance.

Isotopes Mass (amu)	Percentage	abundance
107Ag	106.90509	51.84
109Ag	108.90476	48.16

$$\text{average atomic mass} = \frac{(106.90509 \times 51.84) + (108.90476 \times 48.16)}{100} = 107.87 \text{ amu}$$

Q.7 Boron with atomic number 5 has two naturally occurring isotopes. Calculate the percentage abundance of B<sup>10</sup> and B<sup>11</sup> from the following informations. (Home work)

Average atomic mass of boron = 10.81 amu

Isotopic mass of B<sup>10</sup> = 10.0129 amu

Isotopic mass of B<sup>11</sup> = 11.0093 amu

$$\text{Average atomic mass} = \frac{[10.0129 \times x] + [11.0093 \times (100 - x)]}{100} = 10.81 \text{ amu}$$

$$\text{Or } 10.0129x + 1100.93 - 11.0093x = 1081$$

$$\text{Or } -0.9964x = 1081 - 1100.93$$

$$-0.9964x = -19.93$$

$$x = \frac{19.9}{0.9964} = 20.002\%$$

$$\text{Hence \% abundance of } ^{10}\text{B} = x = 20.002\%$$

$$\text{And \% abundance of } ^{11}\text{B} = 100 - 20.002 = 79.998\%$$

Q.10 Calculate each of the following quantities.

a) Mass in grams of 2.74 moles of KMnO<sub>4</sub>.

(Class work)

Sol:

$$\text{Moles of KMnO}_4 = 2.74 \text{ mol}$$

$$\text{Molar mass of KMnO}_4 = 39 + 55 + 16 \times 4 = 158 \text{ mol}^{-1}$$

$$1 \text{ mole of KMnO}_4 = 158 \text{ g}$$

$$2.74 \text{ mole of KMnO}_4 = 2.74 \times 158 \text{ g} = 432.92 \text{ g}$$

b) Moles of O atoms in 9.00g of Mg (NO<sub>3</sub>)<sub>2</sub>.

(Class work)

Sol:

$$\text{Mass of Mg (NO}_3)_2 = 9 \text{ g}$$

$$\text{Molar mass of Mg (NO}_3)_2 = 24 + 14 \times 2 + 16 \times 6 = 148 \text{ mol}^{-1}$$

$$\text{Moles of Mg (NO}_3)_2 = \frac{\text{Given mass (in grams)}}{\text{Molar mass}}$$

$$= \frac{9}{148} = 0.0608 \text{ moles}$$

$$1 \text{ mole of Mg (NO}_3)_2 = 6 \text{ moles}$$

$$0.0608 \text{ moles contain O moles} = 6 \times 0.0608 = 0.3648 \text{ moles}$$

c) Number of O atoms in 10.037 g of CuSO<sub>4</sub>·5H<sub>2</sub>O. (Class work)

Sol:

$$\text{Mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 10.037 \text{ g}$$

$$\text{Molar mass of CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.5 + 32 + 1 \times 10 + 16 \times 9 = 249.5 \text{ mol}^{-1}$$

$$\text{Formula units of CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \times N_A$$

$$= \frac{10.037}{249.5} \times 6.02 \times 10^{23} = 2.4 \times 10^{22} \text{ formula units}$$

$$1 \text{ Formula units of CuSO}_4 \cdot 5\text{H}_2\text{O} \text{ contain O- atoms} = 9$$

$$2.4 \times 10^{22} \text{ Formula units of CuSO}_4 \cdot 5\text{H}_2\text{O} = 9 \times 2.4 \times 10^{22} = 2.16 \times 10^{23} \text{ atoms}$$

d) Mass in kilograms of  $2.6 \times 10^{20}$  molecules of SO<sub>2</sub>. (Class work)

Sol:

$$\text{Molecules of SO}_2 = 2.6 \times 10^{20}$$

$$\text{Molar mass of SO}_2 = 32 + 16 \times 2 = 64 \text{ gmol}^{-1}$$

$$\text{Molecules of SO}_2 = \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \times N_A$$

$$2.6 \times 10^{20} = \frac{\text{Mass}}{64 \text{ g}} \times 6.02 \times 10^{23}$$



$$\text{Mass} = \frac{2.6 \times 10^{20} \times 64}{6.02 \times 10^{23}} = 2.76 \times 10^{-2} \text{ g}$$

$$\text{Mass in kg} = \frac{2.76 \times 10^{-2}}{1000} = 2.76 \times 10^{-5} \text{ kg}$$

(Class work)

e) Moles of Cl atoms in 0.822 g  $\text{C}_2\text{H}_4\text{Cl}_2$ .

$$\begin{aligned} \text{Sol:} \quad \text{Mass of } \text{C}_2\text{H}_4\text{Cl}_2 &= 0.822 \text{ g} \\ \text{Molar mass of } \text{C}_2\text{H}_4\text{Cl}_2 &= 12 \times 2 + 1 \times 4 + 35.5 \times 2 = 99 \text{ mol}^{-1} \\ \text{Moles of } \text{C}_2\text{H}_4\text{Cl}_2 &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \\ &= \frac{0.822}{99} = 8.303 \times 10^{-3} \text{ moles} \end{aligned}$$

$$\begin{aligned} 1 \text{ mole of } \text{C}_2\text{H}_4\text{Cl}_2 \text{ contain Cl atoms} &= 2 \text{ moles} \\ 8.303 \times 10^{-3} \text{ moles contain Cl atoms} &= 2 \times 8.303 \times 10^{-3} = 0.0166 \text{ moles} \end{aligned}$$

(Class work)

f) Mass in grams of 5.136 moles of  $\text{Ag}_2\text{CO}_3$ .

$$\begin{aligned} \text{Sol:} \quad \text{Mass of } \text{Ag}_2\text{CO}_3 &= 5.136 \text{ mol} \\ \text{Molar mass of } \text{Ag}_2\text{CO}_3 &= 275.74 \text{ g mol}^{-1} \\ \text{Moles of } \text{Ag}_2\text{CO}_3 &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \end{aligned}$$

$$5.136 = \frac{\text{mass}}{275.74}$$

$$\text{Mass} = 5.136 \times 275.74 = 1416.2 \text{ g}$$

g) Mass in grams of  $2.78 \times 10^{21}$  molecules of  $\text{CrO}_2\text{Cl}_2$ .

(Class work)

$$\begin{aligned} \text{Sol:} \quad \text{Molecules of } \text{CrO}_2\text{Cl}_2 &= 2.78 \times 10^{21} \\ \text{Molar mass of } \text{CrO}_2\text{Cl}_2 &= 52 + 16 \times 2 + 35.5 \times 2 = 155 \text{ g mol}^{-1} \\ \text{Molecules of } \text{CrO}_2\text{Cl}_2 &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \times N_A \\ 2.78 \times 10^{21} &= \frac{\text{Mass}}{155} \times 6.02 \times 10^{23} \end{aligned}$$

$$\text{Mass} = \frac{2.78 \times 10^{21} \times 155}{6.02 \times 10^{23}} = 0.7158 \text{ g}$$

h) Number of moles and formula units in 100g of  $\text{KClO}_3$ .

(Class work)

$$\begin{aligned} \text{Sol:} \quad \text{Mass of } \text{KClO}_3 &= 100 \text{ g} \\ \text{Molar mass of } \text{KClO}_3 &= 39 + 35.5 + 16 \times 3 = 122.5 \text{ g mol}^{-1} \\ \text{Moles of } \text{KClO}_3 &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \\ &= \frac{100}{122.5} = 0.816 \text{ moles} \\ \text{Number of Formula units} &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \times N_A \\ &= \frac{100}{122.5} \times 6.02 \times 10^{23} = 4.91 \times 10^{23} \end{aligned}$$

i) Number of  $\text{K}^+$  ions,  $\text{ClO}_3^-$  ions, Cl atoms, and O atoms in (h).

(Class work)

Sol:

$$\text{Formula unit of } \text{KClO}_3 = 4.91 \times 10^{23}$$

$$1 \text{ formula unit of } \text{KClO}_3 \text{ contain } \text{K}^+ \text{ ions} = 1$$

$$4.91 \times 10^{23} \text{ formula unit of } \text{KClO}_3 \text{ contain } \text{K}^+ \text{ ions} = 4.91 \times 10^{23} \text{ ions}$$

$$1 \text{ formula unit of } \text{KClO}_3 \text{ contain } \text{ClO}_3^- \text{ ions} = 1$$

$$4.91 \times 10^{23} \text{ formula unit of } \text{KClO}_3 \text{ contain } \text{ClO}_3^- \text{ ions} = 4.91 \times 10^{23} \text{ ions}$$

$$1 \text{ formula unit of } \text{KClO}_3 \text{ contain Cl atoms} = 1$$

$$4.91 \times 10^{23} \text{ formula unit of } \text{KClO}_3 \text{ contain Cl atoms} = 4.91 \times 10^{23} \text{ atoms}$$

$$1 \text{ formula unit of } \text{KClO}_3 \text{ contain O atoms} = 3$$

$$4.91 \times 10^{23} \text{ formula unit of } \text{KClO}_3 \text{ contain O atoms} = 3 \times 4.91 \times 10^{23} \text{ atoms} = 1.473 \times 10^{24} \text{ atoms}$$

Q.11 Aspartame, the artificial sweetener, has a molecular formula of  $\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$ .

(Class work)

a) What is the mass of one mole of aspartame?

$$\begin{aligned} \text{Sol:} \quad \text{Mass of one mole of aspartame} &= (12 \times 14) + (1 \times 18) + (14 \times 2) + (16 \times 5) \\ &= 294 \text{ g} \end{aligned}$$

b) How many moles are present in 52 g of aspartame?

$$\begin{aligned} \text{Sol:} \quad \text{Moles} &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \\ &= \frac{52}{294} = 0.177 \text{ moles} \end{aligned}$$

c) What is the mass in grams of 10.122 moles of aspartame?

Sol:

$$\text{Moles} = \frac{\text{Given mass (in grams)}}{\text{Molar mass}}$$

$$10.122 \text{ mol} = \frac{\text{Mass}}{294}$$

$$\text{Mass} = 10.122 \times 294 = 2975.87 \text{ g}$$

d) How many hydrogen atoms are present in 2.43 g of aspartame?

$$\begin{aligned} \text{Sol:} \quad \text{Mass of aspartame} &= 2.43 \text{ g} \\ \text{No. of molecules} &= \frac{\text{Given mass (in grams)}}{\text{Molar mass}} \times N_A \\ &= \frac{2.43}{294} \times 6.02 \times 10^{23} = 4.97 \times 10^{21} \text{ molecules} \end{aligned}$$

$$1 \text{ molecule of } \text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5 \text{ contain H atoms} = 18 \text{ atoms}$$

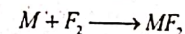
$$4.97 \times 10^{21} \text{ molecules contain H atoms} = 18 \times 4.97 \times 10^{21} = 8.946 \times 10^{22} \text{ atoms}$$

Q.12 A sample of 0.600 moles of a metal M reacts completely with excess of Fluorine to form 46.8 g of  $\text{MF}_2$ .

(Class work)

a) How many moles of F are present in the sample of  $\text{MF}_2$  that forms?

$$\begin{aligned} \text{Sol:} \quad \text{Moles of M} &= 0.600 \text{ mol} \\ \text{Mass of } \text{MF}_2 &= 46.8 \text{ g} \end{aligned}$$





Compare the moles of M and  $\text{MF}_2$  to find the moles of  $\text{MF}_2$

M	:	$\text{MF}_2$
1 mole	:	1 mole
0.600 moles	:	0.600 moles
Thus moles of $\text{MF}_2$ produced	=	0.600 moles
1 moles of $\text{MF}_2$ contain F moles	=	2 moles
0.600 moles of $\text{MF}_2$ contain F moles	=	$2 \times 0.600 = 1.2$ moles

b) Which element is represented by the symbol M?

Sol: Mass of $\text{MF}_2$ produced	=	46.86 g
Moles of $\text{MF}_2$ produced	=	0.600 moles
0.600 moles of $\text{MF}_2$	=	46.8 g
1 mole of $\text{MF}_2$	=	$\frac{46.8}{0.600} = 78$ g
Formula mas of $\text{MF}_2$	=	78
Atomic mas of M + (19 x 2)	=	78
Atomic mass of M	=	$78 - 38 = 40$

Since atomic mass of Ca is 40 therefore M is Ca.

Q.13 In each pair, choose the larger of the indicated quantity, or state if the samples are equal. (Class work)

a) Individual particles: 0.4 mole of oxygen molecules or 0.4 mole of oxygen atoms.

Sol: Both have same number of particles.

Since both are equimolar quantities, therefore they have equal number of particles.

i.e.  $0.4 \times 6.02 \times 10^{23} = 2.408 \times 10^{23}$  particles.

b) Mass: 0.4 mole of ozone molecules or 0.4 mole of oxygen atoms.

Sol: For  $\text{O}_3$

Moles of  $\text{O}_3 = 0.4$  moles

Molar mass of  $\text{O}_3 = 48 \text{ g mol}^{-1}$

Mass = moles x molar mass  
 $= 0.4 \times 48 = 19.2 \text{ g}$

For O

Moles of O = 0.4

Atomic mass of O =  $16 \text{ g mol}^{-1}$

Mass = moles x atomic mass  
 $= 0.4 \times 16 = 6.4 \text{ g}$

c) Mass: 0.6 mole of  $\text{C}_2\text{H}_4$  or 0.6 mole of  $\text{I}_2$ .

Sol: Number of moles = 0.6

Mass = moles x molar mass

Molar mass of  $\text{C}_2\text{H}_4 = 28 \text{ g mol}^{-1}$

Putting the values

Mass of  $\text{C}_2\text{H}_4 = 0.6 \times 28$

$= 16.8 \text{ g}$

Number of moles = 0.6

Mass = moles x molar mass

Molar mass of  $\text{I}_2 = 2 \times 126.9 = 253.8 \text{ g mol}^{-1}$

Mass of  $\text{I}_2 = 0.6 \times 253.8$

$= 152.28 \text{ g}$

d) Individual particles: 4.0 g  $\text{N}_2\text{O}_4$  or 3.3 g  $\text{SO}_2$ .

Sol: Mass of  $\text{N}_2\text{O}_4 = 4 \text{ g}$

Molar mass of  $\text{N}_2\text{O}_4 = 28 + 64 = 92 \text{ g mol}^{-1}$

Number of molecules =  $\frac{\text{Given Mass (in grams)}}{\text{molar mass}}$

Number of molecules =  $\frac{4}{92} \times 6.02 \times 10^{23}$

$= 2.62 \times 10^{22}$  molecules

Mass of  $\text{SO}_2 = 3.3 \text{ g}$

Molar mass of  $\text{SO}_2 = 32 + 32 = 64 \text{ g mol}^{-1}$

Number of molecules =  $\frac{\text{Given Mass (in grams)}}{\text{molar mass}}$

Number of molecules =  $\frac{3.3}{64} \times 6.02 \times 10^{23}$

$= 3.1 \times 10^{22}$  molecules

e) Total ions: 2.3 moles of  $\text{NaClO}_3$  or 2.0 moles of  $\text{MgCl}_2$ .

Sol: Moles of  $\text{NaClO}_3 = 2.3$

1 mole of  $\text{NaClO}_3$  contains  $\text{Na}^+$  ions =  $6.02 \times 10^{23}$

2.3 moles will contain  $\text{Na}^+ = 2.3 \times 6.02 \times 10^{23}$   
 $= 13.846 \times 10^{23}$

1 mole of  $\text{NaClO}_3$  contains  $\text{ClO}_3^-$  ions =  $6.02 \times 10^{23}$

2.3 moles will contain  $\text{ClO}_3^- = 13.846 \times 10^{23}$

Total ions =  $\text{Na}^+$  ions +  $\text{ClO}_3^-$  ions

$= 13.826 \times 10^{23} + 13.846 \times 10^{23}$

$= 2.7692 \times 10^{24}$  Ans.

Moles of  $\text{MgCl}_2 = 2$

1 mole of  $\text{MgCl}_2$  contains  $\text{Mg}^{2+}$  ions =  $6.02 \times 10^{23}$

2 moles will contain  $\text{Mg}^{2+}$  ions =  $6.02 \times 10^{23} \times 2 = 12.04 \times 10^{23}$

1 mole of  $\text{MgCl}_2$  contains  $\text{Cl}^-$  ions =  $2 \times 6.02 \times 10^{23}$

2 moles will contain  $\text{Cl}^-$  ions =  $2 \times 2 \times 6.02 \times 10^{23}$   
 $= 24.08 \times 10^{23}$

Total ions =  $\text{Mg}^{2+}$  ions +  $\text{Cl}^-$  ions =

$= 12.04 \times 10^{23} + 24.08 \times 10^{23}$

$= 3.612 \times 10^{24}$  ions

f) Molecules: 11.0 g  $\text{H}_2\text{O}$  or 11.0 g  $\text{H}_2\text{O}_2$ .

Sol: Mass of  $\text{H}_2\text{O} = 11.0 \text{ g}$

Molar mass of  $\text{H}_2\text{O} = 18 \text{ g mol}^{-1}$

Number of molecules =  $\frac{\text{Given Mass (in grams)}}{\text{molar mass}}$

Number of molecules of  $\text{H}_2\text{O} = \frac{11}{18} \times 6.02 \times 10^{23}$   
 $= 3.67 \times 10^{23}$

Mass of  $\text{H}_2\text{O}_2 = 11.0 \text{ g}$

Molar mass of  $\text{H}_2\text{O}_2 = 34 \text{ g mol}^{-1}$

Number of molecules =  $\frac{\text{Given Mass (in grams)}}{\text{molar mass}}$



$$\text{Number of molecules of } H_2O = \frac{11}{34} \times 6.02 \times 10^{23} = 1.926 \times 10^{23}$$

g)  $Na^+$  ion: 0.500 moles of NaBr or 0.0145 kg of NaCl.

Sol: Moles of NaBr = 0.500

0.500 moles of NaBr contain 0.5 mole of  $Na^+$  ions

Number of ions = moles  $\times N_A$

$$= 0.5 \times 6.02 \times 10^{23}$$

$$= 3.01 \times 10^{23} \text{ ions}$$

$$\text{Mass of NaCl} = 0.0145 \text{ kg} = 14.5 \text{ g}$$

$$\text{Molar mass of NaCl} = 58.5 \text{ g mol}^{-1}$$

$$\text{Moles of NaCl} = \frac{\text{Give Mass (in grams)}}{\text{Molar mass}} = \frac{14.5}{58.5} = 0.2478 \text{ moles}$$

0.2478 moles of NaCl contain 0.2478 moles of

$Na^+$  ions, therefore

$$\text{Number of } Na^+ \text{ ions} = \text{moles} \times N_A$$

$$= 0.2478 \times 6.02 \times 10^{23}$$

$$= 1.49 \times 10^{23}$$

h) Mass:  $6.02 \times 10^{23}$  atoms of  $^{235}U$  or  $6.02 \times 10^{23}$  atoms of  $^{238}U$ .

Sol:  $6.02 \times 10^{23}$  atoms means 1 mole. Therefore

$$6.02 \times 10^{23} \text{ atoms of } ^{235}U = 235 \text{ g}$$

$$6.02 \times 10^{23} \text{ atoms means 1 mole, therefore,}$$

$$6.02 \times 10^{23} \text{ atoms of } ^{238}U = 238 \text{ g}$$

Q.14 a) Calculate the percentage of nitrogen in the four important fertilizers i.e., (i)  $NH_3$  (ii)  $NH_2CONH_2$  (urea) (iii)  $(NH_4)_2SO_4$  (iv)  $NH_4NO_3$ . (Class work)

Sol: % is calculated by the formula

$$\% \text{ of nitrogen} = \frac{\text{Mass of N atoms in the compound}}{\text{Molar mass of the compound}} \times 100$$

$$\text{Molar Mass of } NH_3 = 14 \times 1 + 1 \times 3 = 17 \text{ g mol}^{-1}$$

$$\% \text{ of N in } NH_3 = \frac{14}{17} \times 100 = 82.35\%$$

$$\text{Molar Mass of } NH_2CONH_2 = 14 \times 2 + 1 \times 4 + 12 \times 1 + 16 \times 1 = 60 \text{ g mol}^{-1}$$

$$\% \text{ of N in } NH_2CONH_2 = \frac{28}{60} \times 100 = 46.67\%$$

$$\text{Molar Mass of } (NH_4)_2SO_4 = 14 \times 2 + 1 \times 8 + 32 \times 1 + 16 \times 4 = 132 \text{ g mol}^{-1}$$

$$\% \text{ of N in } (NH_4)_2SO_4 = \frac{28}{132} \times 100 = 21.21\%$$

$$\text{Molar Mass of } NH_4NO_3 = 14 \times 2 + 1 \times 4 + 16 \times 3 = 80 \text{ g mol}^{-1}$$

$$\% \text{ of N in } NH_4NO_3 = \frac{28}{80} \times 100 = 35\%$$

b) Calculate the percentage of nitrogen and phosphorus in each of the following: (i)  $NH_4H_2PO_4$  (ii)  $(NH_4)_2HPO_4$  (iii)  $(NH_4)_3PO_4$

Sol: % of N and P are calculated by the formula

$$\% \text{ of N or P} = \frac{\text{Mass of N or P atoms in the compounds}}{\text{Molar mass of the compound}} \times 100$$

$$(i) \text{ Molar Mass of } NH_4H_2PO_4 = 14 \times 1 + 1 \times 6 + 31 \times 1 + 16 \times 4 = 115 \text{ g mol}^{-1}$$

$$\% \text{ of N in } NH_4H_2PO_4 = \frac{14}{115} \times 100 = 12.17\%$$

$$\% \text{ of P in } NH_4H_2PO_4 = \frac{31}{115} \times 100 = 26.96\%$$

$$(ii) \text{ Molar Mass of } (NH_4)_2HPO_4 = 14 \times 2 + 1 \times 9 + 31 \times 1 + 16 \times 4 = 132 \text{ g mol}^{-1}$$

$$\% \text{ of N in } (NH_4)_2HPO_4 = \frac{31}{132} \times 100 = 23.35\%$$

$$\% \text{ of P in } (NH_4)_2HPO_4 = \frac{31}{115} \times 100 = 26.96\%$$

$$(iii) \text{ Molar Mass of } (NH_4)_3PO_4 = 14 \times 3 + 1 \times 12 + 31 \times 1 + 16 \times 4 = 149 \text{ g mol}^{-1}$$

$$\% \text{ of N in } (NH_4)_3PO_4 = \frac{42}{149} \times 100 = 28.19\%$$

$$\% \text{ of P in } (NH_4)_3PO_4 = \frac{31}{149} \times 100 = 20.80\%$$

Q.15 Glucose  $C_6H_{12}O_6$  is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass % of each element in glucose and determine the number of C, H and O atoms in 10.5 g of the sample. (Class work)

Sol: Molar Mass of  $C_6H_{12}O_6 = 12 \times 6 + 1 \times 12 + 16 \times 6 = 180 \text{ g/mol}$

$$\% \text{ of an element} = \frac{\text{Mass of the element in the compounds}}{\text{Molar mass of the compound}} \times 100$$

$$\% \text{ of carbon} = \frac{72}{180} \times 100 = 40\%$$

$$\% \text{ age of H} = \frac{12}{180} \times 100 = 6.67\%$$

$$\% \text{ age of O} = \frac{96}{180} \times 100 = 53.33\%$$

$$\text{Mass of } C_6H_{12}O_6 = 10.5 \text{ g}$$

$$\text{Molecules of } C_6H_{12}O_6 = \frac{\text{Given mass (in gram)}}{\text{Molar mass}} \times N_A$$

$$= \frac{10.5}{180} \times 6.02 \times 10^{23} = 3.5 \times 10^{22} \text{ molecule}$$

$$1 \text{ molecule of } C_6H_{12}O_6 \text{ contains C atoms} = 6 \text{ atoms}$$

$$3.5 \times 10^{22} \text{ molecule of } C_6H_{12}O_6 \text{ contains C atoms} = 6 \times 3.5 \times 10^{22} = 2.1 \times 10^{23} \text{ atoms}$$

$$1 \text{ molecule of } C_6H_{12}O_6 \text{ contains H atoms} = 12 \text{ atoms}$$

$$3.5 \times 10^{22} \text{ molecule of } C_6H_{12}O_6 \text{ contains H atoms} = 12 \times 3.5 \times 10^{22} = 4.2 \times 10^{23} \text{ atoms}$$

$$1 \text{ molecule of } C_6H_{12}O_6 \text{ contains O atoms} = 6 \text{ atoms}$$

$$3.5 \times 10^{22} \text{ molecule of } C_6H_{12}O_6 \text{ contains O atoms} = 6 \times 3.5 \times 10^{22} = 2.1 \times 10^{23} \text{ atoms}$$

Q.16 Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7 % hydrogen and 51.6% oxygen. Its molar mass is 62.1 grams mol<sup>-1</sup>. Determine its empirical formula? (Class work)

Sol:



A Plus Chemistry (ALP Smart Syllabus-2020)

11<sup>th</sup> Class

ELEMENT	%	No. of gram Atoms	Atomic Ratio	Empirical Formula
C	38.7	$\frac{38.7}{12} = 3.225$	$\frac{3.225}{3.225} = 1$	CH <sub>3</sub> O
H	9.7	$\frac{9.7}{1.008} = 9.6$	$\frac{9.6}{3.225} = 3$	
O	51.6	$\frac{51.6}{16} = 3.225$	$\frac{3.225}{3.225} = 1$	

Thus Empirical Formula of Etylene Glycol is CH<sub>3</sub>O

$$\text{Empirical Formula mass} = (12 \times 1) + (3 \times 1) + (16 \times 1) = 31 \text{ g mol}^{-1}$$

$$\text{Molecular mass} = 62.1 \text{ g mol}^{-1}$$

$$n = \frac{\text{Molecular Mass}}{\text{Empirical Mass}} = \frac{62.1}{31} = 2$$

$$\text{Molecular Formula} = n(\text{Empirical Formula}) = 2(\text{CH}_3\text{O}) = \text{C}_2\text{H}_6\text{O}_2$$

Q.17 Serotonin (Molar mass = 176 g mol<sup>-1</sup>) is a compound that conducts nerve impulse in brain and muscles. It contains 68.2 % C, 6.86 % H, 15.09 % N, and 9.08 % O. What is its molecular formula. (Class work)

Sol:

ELEMENT	%	No. of gram Atoms	Atomic Ratio	Empirical Formula
C	68.20	$\frac{68.20}{12} = 5.7$	$\frac{5.7}{0.56} = 10$	C <sub>10</sub> H <sub>12</sub> N <sub>2</sub> O
H	6.86	$\frac{6.86}{1.008} = 6.8$	$\frac{6.8}{0.56} = 12$	
N	15.09	$\frac{15.09}{14} = 1.07$	$\frac{1.07}{0.56} = 2$	
O	9.08	$\frac{9.08}{16} = 0.56$	$\frac{0.56}{0.56} = 1$	

Thus Empirical Formula of Etylene Glycol is C<sub>10</sub>H<sub>12</sub>N<sub>2</sub>O

$$\text{Empirical Formula mass} = (12 \times 10) + (1 \times 12) + (14 \times 2) + (16 \times 1) = 176 \text{ g mol}^{-1}$$

$$\text{Molecular mass} = 176 \text{ g mol}^{-1}$$

$$n = \frac{\text{Molecular Mass}}{\text{Empirical Mass}} = \frac{176}{176} = 1$$

$$\text{Molecular Formula} = n(\text{Empirical Formula}) = 1 \text{ C}_{10}\text{H}_{12}\text{N}_2\text{O} = \text{C}_{10}\text{H}_{12}\text{N}_2\text{O}$$

Q.18 An unknown metal M reacts with S to form a compound with a formula M<sub>2</sub>S<sub>3</sub>. 3.12 g of M reacts with exactly 2.88 g of sulphur, what are the names of metal M and the compound M<sub>2</sub>S<sub>3</sub>?

Sol: The reaction can be written as  $2M + 3S \rightarrow M_2S_3$  (Class work)

$$\text{Mass of S in } M_2S_3 = 32 \times 3 = 96 \text{ g}$$

$$\text{According to the given data } 2.88 \text{ g of S react with M} = 3.12 \text{ g}$$

$$96 \text{ g of S react with M} = \frac{3.12 \text{ g}}{2.88 \text{ g}} \times 96 \text{ g} = 104 \text{ g}$$

Thus according to the formula M<sub>2</sub>S<sub>3</sub>

$$2 \text{ moles of M} = 104 \text{ g of M}$$

$$1 \text{ mole of M} = \frac{104}{2} = 52 \text{ g mol}^{-1}$$

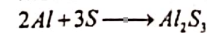
$$\text{Hence Atomic mass of M} = 52 \text{ g mol}^{-1}$$

Since 52 g mol<sup>-1</sup> is the atomic mass of chromium.

Therefore M is chromium and M<sub>2</sub>S<sub>3</sub> is Cr<sub>2</sub>S<sub>3</sub> i.e. chromium sulphide

Q.20 Calculate the number of grams of Al<sub>2</sub>S<sub>3</sub> which can be prepared by the reaction of 20 g of Al and 30 g of sulphur. How much the non-limiting reactant is in excess? (Class work)

Sol: The balanced chemical equation for the chemical reaction is



$$\text{Given mass of Al} = 20 \text{ g}$$

$$\text{Moles of Al} = \frac{20}{27} = 0.74 \text{ moles}$$

$$\text{Given mass of S} = 30 \text{ g}$$

$$\text{Moles of S} = \frac{30}{32} = 0.9375 \text{ moles}$$

Determination of Limiting Reactant

Compare of Al and Al<sub>2</sub>S<sub>3</sub> according to balanced chemical equation.

$$\begin{array}{lcl} \text{Al} & : & \text{Al}_2\text{S}_3 \\ 2 \text{ moles} & : & 1 \text{ mole} \end{array}$$

$$\text{Therefore } 0.74 \text{ moles} : \frac{1}{2} \times 0.74 = 0.37 \text{ moles}$$

Hence Moles of the product produced by Al = 0.37 moles

Compare S and Al<sub>2</sub>S<sub>3</sub> according to balanced chemical equation.

$$\begin{array}{lcl} \text{S} & : & \text{Al}_2\text{S}_3 \\ 3 \text{ moles} & : & 1 \text{ mole} \end{array}$$

$$\text{Therefore } 0.9375 \text{ moles} : \frac{1}{3} \times 0.9375 = 0.3125 \text{ moles}$$

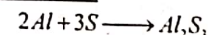
Hence Moles of the Al<sub>2</sub>S<sub>3</sub> produced by S = 0.3125 moles

Since S produces least number of moles of the product, therefore it is the limiting reactant.

$$\text{Moles of the Al}_2\text{S}_3 \text{ produced} = 0.3125 \text{ moles}$$

$$\text{Mass of Al}_2\text{S}_3 \text{ produced} = 0.3125 \times 150 = 46.87 \text{ g}$$

Determination of amount of Al left unreacted



Compare the moles of S and Al to find the moles of Al reacted.

$$\begin{array}{lcl} \text{S} & : & \text{Al} \\ 3 \text{ moles} & : & 2 \text{ mole} \end{array}$$

$$0.9375 \text{ moles} : \frac{2}{3} \times 0.9375 = 0.625 \text{ moles}$$

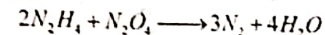
$$\text{Moles of Al consumed} = 0.625 \text{ moles}$$

$$\text{Moles of Al taken} = 0.74 \text{ moles}$$

$$\text{Moles of Al left unreacted} = 0.74 - 0.625 = 0.115 \text{ moles}$$

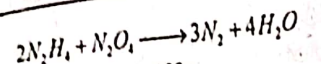
$$\text{Mass of Al left unreacted} = 0.115 \times 27 = 3.105 \text{ g}$$

Q.21 A mixture of two liquids, hydrazine N<sub>2</sub>H<sub>4</sub> and N<sub>2</sub>O<sub>4</sub> are used in rockets. They produce N<sub>2</sub> and water vapours. How many grams of N<sub>2</sub> gas will be formed by reacting 100g of N<sub>2</sub>H<sub>4</sub> and 200g of N<sub>2</sub>O<sub>4</sub>. (Class work)



Sol: The balanced chemical equation for the chemical reaction is





Given mass of $N_2H_4$	=	100g
Moles of $N_2H_4$	=	$\frac{100}{32} = 3.125 \text{ moles}$
Given mass of $N_2O_4$	=	200g
Moles of $N_2O_4$	=	$\frac{200}{92} = 2.17 \text{ moles}$

**Determination of Limiting Reactant**

Compare of  $N_2H_4$  and  $N_2O_4$  according to balanced chemical equation.

$N_2H_4$	:	$N_2$
2 moles	:	3 mole
3.125 moles	:	$\frac{3}{2} \times 3.125 = 4.69 \text{ moles}$

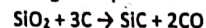
Compare the moles of  $N_2O_4$  and  $N_2$  to find the moles of  $N_2$  reacted.

$N_2O_4$	:	$N_2$
1 mole	:	3 mole
Therefore 2.17 moles	:	$\frac{3}{1} \times 2.17 = 6.51 \text{ moles}$

Since  $N_2H_4$  produce least number of moles of the product, therefore it is the limiting reactant.

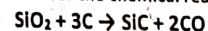
Hence Moles of the product produced	=	4.69 moles
Mass of the product produced	=	$4.67 \times 28 = 131.32 \text{ g}$

Q.22 Silicon carbide (SiC) is an important ceramic material. It is produced by allowing sand ( $SiO_2$ ) to react with carbon at high temperature.



When 100 kg sand is reacted with excess of carbon, 51.4 kg of SiC is produced. What is the percentage yield of SiC?

Sol: The balanced chemical equation for the chemical reaction is:



Mass of $SiO_2$	=	100 kg = 100000g
Mole of $SiO_2$	=	$\frac{100000}{60} = 1666.67 \text{ moles}$

Compare  $SiO_2$  and SiC according to balanced chemical equation.

$SiO_2$	:	SiC
1 mole	:	1 mole
Therefore 1666.67 moles	:	1666.67 moles

Thus

Moles of SiC, expected to be produced	=	1666.67 moles
Molar mass of SiC	=	$28 + 12 = 40 \text{ g/mol}$
Mass of SiC expected to be produced	=	$1666.67 \times 40 = 66666.8 \text{ g}$
Theoretical yield of SiC	=	66666.8g
Actual yield of SiC	=	51.4 kg = 51400g

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

$$= \frac{51400}{66666.8} \times 100 = 77\%$$

## CHAPTER-2 (EXPERIMENTAL TECHNIQUES)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS 2020)

- Silica gel can be used as  
(a) Oxidizing agent (b) dehydrating agent (c) Hydrating agent (d) reducing agent
- The most common solvent used in solvent extraction is:  
(a) Acetone (b) Ethanol (c) Rectified spirit (d) Diethyl ether
- Solvent extraction is a process:  
(a) Exothermic (b) Endothermic (c) Equilibrium (d) Non-equilibrium
- Iodine dissolve in water in the presence of KI due to formation of which one of the following species:  
(a)  $I_2$  (b)  $I^{(-)}$  (c)  $I_3^{(-)}$  (d)  $I_4$  (1 time)
- During chromatography strip should be dipped in to solvent mixture to a depth of:  
(a) 3-4mm (b) 4-5mm (c) 5-6mm (d) 6-7mm
- The chromatography in which stationary phase is liquid is called:  
(a) thin layer chromatography (b) partition chromatography  
(c) absorption chromatography (d) Gel chromatography

**Answers**

1	2	3	4	5	6
b	d	c	c	c	b

## CHAPTER-2 (EXPERIMENTAL TECHNIQUES)

### SHORT QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

- Give the main characteristics of the solvent used for crystallization? (6 times)  
(Home work)  
Ans: i) It should not react chemically with solute. ii) It should be cheap. iii) It should be safe to use. iv) It should be easily removable.
- State solvent extraction and give its importance? (2 times)  
Ans: Solvent extraction: The gain of a substance from a solution with the help of an immiscible solvent is called solvent extraction. This method of separation is mostly used to separate organic compounds from water.
- Sublimation:**
- How mixture of  $NH_4Cl$  and  $NaCl$  can be separated? (2 times)  
Ans: Mixture of  $NH_4Cl$  and  $NaCl$  can be separated by sublimation process, because ammonium chloride sublime while  $NaCl$  can not.
- Define sublimation with two examples? (9 times)  
Ans: Sublimation: The process in which a solid is directly converted into vapors without passing through the liquid phase, and then it can be condensed to obtain solid again.  
e.g. ammonium chloride, iodine, benzoic acid.
- Define sublimation and partition law? (2 times)  
Ans: Sublimation is a process in which a solid, when heated vapourizes directly without passing through the liquid state.  
Partition law states that a solute distributes itself between two immiscible liquids in a constant ratio of concentration irrespective of the amount of solute added.
- How purification is carried out by Sublimation?  
Ans: Sublimation is a process in which a solid, when heated, vapourizes directly without passing through the liquid phase and these vapours can be condensed to form the solid again.  
To carry out the process, the substance is taken in a watch glass covered with an inverted funnel. The substance is then heated slowly over a sand-bath and the



funnel is cooled with wet cotton.

### Solvent Extraction:

4. Iodine is more soluble in water in the presence of KI. Discuss? (13 times)
- Ans: As we know that like dissolve like so KI when dissolve in  $H_2O$  give iodide ion which is soluble in  $H_2O$ . But  $I_2$  is a non-polar compound and is not dissolved in  $H_2O$ . Iodine in presence of  $I^-$  ion is converted to  $I_3^-$ , which is water soluble. (12 times)
8. State distribution law or partition law?
- Ans: Distribution law: At constant temperature a solute distributes itself between two immiscible liquids in a constant ratio of concentration independent of the amount of solute added. (2 times)
9. Define distribution law and partition law?
- Ans: Solvent extraction is an equilibrium process and follows the distribution law or partition law. This law states that a solute distributes itself between two immiscible liquids in a constant ratio of concentration irrespective of the amount of solute added. (2 times)
10. Define Solvent Extraction and Partition Law?
- Ans: Solvent extraction is a technique in which a solute can be separated from a solution by shaking the solution with a solvent in which the solute is more soluble and the added solvent does not mix with the solution. Solvent extraction is an equilibrium process and follows the distribution law or partition law. This law states that a solute distributes itself between two immiscible liquids in a constant ratio of concentration irrespective of the amount of solute added. (2 times)
11. Describe Ether Extraction.
- Ans: The most common laboratory example of solvent extraction is ether extraction. This is used to separate the products of organic synthesis from water. In a typical organic synthesis, the aqueous solution containing the organic product is shaken up with ether in a separating funnel and allowed to separate. The inorganic impurities remain in aqueous phase, whereas the organic compound goes to the ether layer. The ether layer is separated and organic product is obtained by evaporating the ether. (2 times)
12. What do you mean by solvent extraction? Which law controls it?
- Ans: Solvent extraction is a technique in which solute can be separated from a solution by shaking the solution with a solvent in which the solute is more soluble and the added solvent does not mix with the solution. Usually it is done by placing the solution and the second liquid into a separating funnel. The funnel is stoppered and the two liquids are shaken together. This process is controlled by distribution law. (2 times)
13. State Law of distribution constant.
- Ans: Solvent extraction is done by placing the solution and the second liquid into a separating funnel. The funnel is stoppered and the two liquids are shaken together. This process is controlled by distribution law. The ratio of the final concentration at equilibrium is constant. This is called distribution coefficient, K. In case of distribution of iodine in  $CCl_4$ , K is given by:
- $$K = \frac{[I_2(CCl_4)]}{[I_2]}$$

### 5. Chromatography:

14. What do you mean by (a) chromatogram (b) Filtrate?
- Ans: (a) Chromatogram: The dried paper at which various components have been separated is called chromatogram. From chromatogram  $R_f$  values are determined. (b) Filtrate: The liquid which is obtained after passing through the filter medium is called filtrate. (2 times)
15. Differentiate between stationary phase and mobile phase?
- Ans: Stationary phase: It may be a solid or liquid supported by an inert solid. Mobile phase: It may be a liquid or a gas. It flows over stationary phase. (2 times)
16. Define chromatography and give formula of distribution coefficient? (3 times)
- Ans: It is an analytical technique used for the separation of a mixture due to different distribution of substances between stationary and mobile phase.
- $$K = \frac{\text{Concentration of a component in stationary phase}}{\text{Concentration of a component in mobile phase}}$$
17. Differentiate between adsorption and partition chromatography? (10 times)
- Ans: Adsorption Chromatography: Chromatography in which stationary phase is a solid is called adsorption chromatography. e.g. column chromatography. (10 times)

**Partition Chromatography:** Chromatography in which stationary phase is a liquid is called partition chromatography. e.g. paper chromatography.

18. In solvent extraction technique, repeated extraction using small portions of solvent are more efficient than using a single extraction, but larger volume of solvent?

Ans: Separation of solute depend upon the distribution ratio of solute in mobile and stationary phase  $K = \frac{C_m}{C_s}$

Repeated extraction with small portion of solvents give better separation because again and again equilibrium is established between two phases. More amount of solute is distributed among the mobile phase using small portion.

19. Define  $R_f$  value and why it has no unit? (6 times)

Ans:  $R_f$  value:  $R_f$  value is defined as the ratio of the distance travelled by a component from a base line to the distance travelled by solvent from base line.

$$R_f = \frac{\text{Distance travelled by a component from base line}}{\text{Distance travelled by solvent from base line}}$$

It has no units because it is ratio of distance to distance.

20. Give two application of paper chromatography?

Ans: i) It is used for separation and purification of substances from a mixture.

ii) It is used for qualitative and quantitative analysis.

21. Give two uses of chromatography? (3 times)

Ans: The techniques of chromatography are very useful in organic synthesis for separation, isolation and purification of the products. They are equally important in qualitative and quantitative analysis and for determination of the purity of a substance.

22. What is solvent extraction. Give its importance.

Ans: Solvent extraction is a technique in which a solute can be separated from a solution by shaking the solution with a solvent in which solute is more soluble and added solvent does not mix with the solution.

Solvent extraction is an important technique of chemical analysis. It is used when substance to be separated is volatile or thermally unstable.

23. Write names of substances purified by sublimation.

Ans: (1) Benzoic acid (2) Ammonium chloride  
(3) Iodine (4) Naphthalene

## Solved Exercise Chapter # 2 ALP SMART SYLLABUS 2020

- Q.1 Multiple choice questions.

(Class work)

(iii) Solvent extraction is an equilibrium process and it is controlled by.

- (a) law of mass action. (b) the amount of solvent used.  
(c) distribution law. (d) the amount of solute.

(iv) Solvent extraction method is a particularly useful technique for separation when the product to be separated is:

- (a) non-volatile or thermally unstable. (b) volatile or thermally stable.  
(c) non-volatile or thermally stable. (d) volatile or thermally unstable.

(v) The comparative rates at which the solutes move in paper chromatography, depend on:

- (a) the size of paper (b)  $R_f$  values of solutes.  
(c) temperature of the experiment. (d) size of the chromatographic tank used.

2- Fill in the blanks. (Class work)

i. A complete chemical characterization of a compound must include qualitative and quantitative analysis.

iv. A solvent used for crystallization is required to dissolve large amount of the substance at its boiling point and small amount at the room temperature.

v. Repeated solvent extractions using small portions of solvent are more efficient than using a single extraction with larger volume of the solvent.



11<sup>th</sup> Class

28

Atomic

Q.3 Tick the correct sentences. If the sentence is incorrect, write the correct statements.

(i) A solute distributes itself between two immiscible liquids in a constant ratio of concentrations depending upon the amount of solvent added. (False)

(ii) Paper chromatography is a technique of partition chromatography. (True)

(iii) A solid organic compound is soluble in water as well as in chloroform. During its preparation, it remains in aqueous layer. Describe a method to obtain from this layer. (Home work)

(iv) A compound can be obtained by solvent extraction technique. In the aqueous layer. Since given organic compound is soluble in organic solvent.

Q.6. The organic compound can be obtained by solvent extraction technique. In the given conditions, the organic compound is present in aqueous layer. Since given organic compound is also soluble in chloroform and chloroform is almost immiscible with  $H_2O$ . Hence, organic compound can be separated by shaking aqueous layer with chloroform in a separating funnel. The compound goes into the chloroform layer, which is separated from aqueous layer. Then chloroform is evaporated to get pure organic compound.

Q.7. The following figure shows a developed chromatogram on paper with five spots.

Q.7 The following figure shows a developed chromatogram.

(i) Unknown mixture X	(ii) Sample A	(v) Sample D
(iii) Sample B	(iv) Sample C	

Find out (i) the composition of unknown mixture X

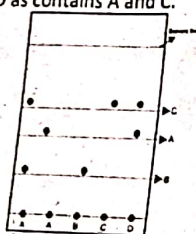
(ii) which sample is impure and what is its composition.

**Sol:** (i) the composition of unknown mixture X

1) the composition of unknown mixture X  
The chromatogram shows that the mixture X contains component B and C.

(ii) which sample is impure and what is its composition.

The impure sample is D as contains A and C.



Q.8 In solvent extraction technique, why repeated extraction using small portions of solvent are more efficient than using a single extraction but larger volume of solvent?

Sol: Repeated extraction using small portions of solvent are more efficient because more amount of substance is extracted.

Q.10 You have been

Sol: Repeated extraction using small portions of solvent is more efficient because more amount of substance is extracted.

**Q.10** You have been provided with a mixture containing three inks with different colours. Write down the procedure to separate the mixture with the help of paper chromatography.

**Sol:** Inks can be separated by following method.

- A solvent mixture

(Home work)

**Sol:** Inks can be separated by following method.

- A solvent mixture is placed in a chromatographic tank.
- The tank is covered with a lid so that the solvent does not evaporate.

(Home work)

- A solvent mixture is placed in a chromatographic tank.
  - The tank is covered with a lid so that its inner space is saturated with solvent vapours and becomes homogenous.
  - About 20cm strip of Whatmann's Chromatographic paper No.1 is taken.
  - A line is drawn with a thin pencil, about 2.5cm above from one end of paper. This is the base line.
  - A drop of mixture of inks is placed on the base line.
  - After drying the spots, the paper is suspended in the chromatographic tank and left for sufficient time.
  - When solvent front has moved to  $3/4^{\text{th}}$  of the length of the paper, paper is removed from the tank.
  - Solvent front is marked with a pencil line and paper is dried.
  - The spots of inks can be seen on paper.
- For each type of ink  $R_f$  values can be determined.

$R_i = \frac{\text{Distance travelled by a particular ink from base}}{\text{Distance travelled by the solvent front}}$

$R_f = \frac{\text{Distance travelled by a particular ink from base line}}{\text{Distance travelled by the solvent from base line}}$

### CHPATER-3 (GASES ) OBJECTIVES (MCQ'S)

1. Formula used for conversion of  $^{\circ}\text{F}$  into  $^{\circ}\text{C}$  is:  
(a)  $^{\circ}\text{F} = 9/5(^{\circ}\text{C}) + 32$  (b)  $^{\circ}\text{C} = 5/9[^{\circ}\text{F} - 32]$  (c)  $^{\circ}\text{F} = 5/9(^{\circ}\text{C}) + 32$  (d)  $^{\circ}\text{C} = 9/5[^{\circ}\text{F} - 32]$
2. An ideal gas has volume  $1\text{ dm}^3$  at  $303\text{ K}$ . Keeping pressure constant. At which Kelvin temperature its volume will become  $2\text{ dm}^3$ .  
(a)  $240\text{ K}$  (b)  $303\text{ K}$  (c)  $330\text{ K}$  (d)  $606\text{ K}$
3. The pair of gases which does not obey Dalton's law of partial pressure under normal condition is:  
(a)  $\text{H}_2$  and  $\text{O}_2$  (b)  $\text{H}_2$  and  $\text{He}$  (c)  $\text{NH}_3$  and  $\text{HCl}$  (d)  $\text{He}$  and  $\text{Ne}$
4. Partial pressure of oxygen in the air is: (2 times)  
(a)  $156\text{ torr}$  (b)  $157\text{ torr}$  (c)  $158\text{ torr}$  (d)  $159\text{ torr}$
5. Partial pressure of oxygen in human lungs in torr is: (6 times)  
(a)  $161$  (b)  $116$  (c)  $159$  (d)  $760$
6. Normal human body temperature is:  
(a)  $37^{\circ}\text{C}$  (b)  $98.6^{\circ}\text{C}$  (c)  $37^{\circ}\text{F}$  (d)  $273\text{ K}$
7. When water freezes at  $0^{\circ}\text{C}$ , its density decreases due to:  
(a) Empty spaces present in the structure of ice. (b) Cubic structure of ice  
(c) Change of bond lengths (d) Change of bond angles
8. Mass of  $22.4\text{ dm}^3$  of  $\text{N}_2$  at STP is: (3 times)  
(a)  $28\text{ gm}$  (b)  $14\text{ gm}$  (c)  $1.4\text{ gm}$  (d)  $2.8\text{ gm}$
9. In 1879, plasma was identified by scientist: (1 time)  
(a) John Dalton (b) Chadwick (c) William Crookes (d) Soddy
10. Plasma is:  
(a) First state of matter (b) Second state of matter  
(c) Third state of matter (d) Fourth state of matter
11. The molar volume of  $\text{O}_2$  gas is maximum at:  
(a) STP (b)  $127^{\circ}\text{C}$  and  $1\text{ atm}$  (c)  $0.00^{\circ}\text{C}$  and  $2\text{ atm}$  (d)  $273^{\circ}\text{C}$  and  $2\text{ atm}$

## Answers

1	2	3	4	5	6	7	8	9	10	11				
b	d	c	d	b	a	a	a	c	d	b				

## CHAPTER-3 (GASES ) SHORT QUESTION'S (ACCRODING TO ALP SMART SYLLABUS-2020)

### Gas Laws:

1. Define Boyle's law and give its mathematical expression? (3 times) (Home work)
- Ans: Boyle's law: At constant temperature volume of given mass of a gas is inversely proportional to the pressure exerted on it.

$$V \propto \frac{1}{P} \quad (\text{At constant temperature and number of moles})$$

$$V = \frac{K}{P}$$

$$PV = K$$

2. Why do we get a straight line when pressures are plotted against inverse of volume for gas? (Home work)



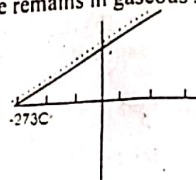
The plot of "P" on y-axis against  $1/V$  on x-axis gives a straight line. It shows that "P" is directly proportional to  $1/V$ . This straight line passes through the origin because when the pressure is close to zero then volume is so high that  $1/V$  is very close to zero.



3. What is absolute zero? What happens to real gases while approaching this temperature? (11 times) (Home work)

Ans: The hypothetical temperature at which the volume of all the gases becomes zero is called absolute zero. In Celsius scale it is  $-273.16^\circ\text{C}$ .  
Graphical expression:

When a graph is plotted between V and T for a gas a straight line is obtained which intersects the temperature axis at  $-273^\circ\text{C}$  which is considered as the lowest temperature. It would be achieved if the substance remains in gaseous state. But all



gases liquefy before reaching this temperature.

4. Convert  $-40^\circ\text{F}$  temperature to (a) Centigrade scale (b) Kelvin scale? (3 times)

Ans: (i) Centigrade scale

$$C^\circ = \frac{5}{9}(F^\circ - 32)$$

$$C^\circ = \frac{5}{9}(-40 - 32) = -40$$

(ii) Kelvin scale

$$K = C^\circ + 273.16$$

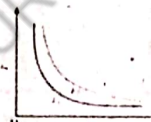
$$K = -40 + 273.16 = 233.16\text{K}$$

5. What is absolute scale?

Ans: Kelvin scale of temperature is called absolute temperature scale because it starts from absolute zero, i.e.  $-273.16^\circ\text{C}$ .

6. What are isotherms? (Home work)

Ans: Isotherms: ('Iso' means same and 'thermo' means heat)  
The plot of "V" and "P" at constant temperature gives a curve called isotherm.



7. Define thermometry. Name the scale and device used for thermometry?

Ans: Thermometry: The study of measurement of temperature is called thermometry. Temperature is the average kinetic energy of particles of a substance. There are three measuring scales of thermometry.

(i) Celsius or Centigrade scale ( $C^\circ$ )

(ii) Fahrenheit scale ( $F^\circ$ )

(iii) Kelvin scale (K)

8. State Charles's Law and write its mathematical form? (2 times)

Ans: The volume of the given mass of gas is directly proportional to the absolute temperature when the pressure is kept constant is known as Charles's law. Mathematically it can be written as:

$$\frac{V}{T} = k \quad (\text{at constant pressure and number of moles})$$

If the temperature is changed from  $T_1$  to  $T_2$  and volume change from  $V_1$  to  $V_2$  then

$$\frac{V_1}{T_1} = k \quad \text{and} \quad \frac{V_2}{T_2} = k$$

9.  $\text{SO}_2$  is non-ideal at  $273\text{K}$ , but behaves like an ideal gas at  $327\text{K}$ . Justify the statement? (5 times)

Ans: At low temperature of  $273\text{K}$  molecules of  $\text{SO}_2$  gas have considerable attractions for each other and thus  $\text{SO}_2$  gas behaves non-ideally. When temperature is increased to  $327\text{K}$  the forces of attraction among  $\text{SO}_2$  molecules decreases and hence  $\text{SO}_2$  gas shows ideal behaviours.

10. Define Quantitative Definition of Charles's Law.

Ans: At constant pressure, the volume of the given mass of gas increases or decreases by  $\frac{1}{273}$  of its original volume at  $0^\circ\text{C}$  for every  $1^\circ\text{C}$  rise or fall in temperature respectively.

11. Throw some light on the factor  $\frac{1}{273}$  in Charles law. (Home work)

Ans: At constant pressure, the volume of the given mass of gas increases or decreases by  $\frac{1}{273}$  of its original volume at  $0^\circ\text{C}$  for every  $1^\circ\text{C}$  rise or fall in temperature respectively.

If a gas is warmed by  $1^\circ\text{C}$ , it expands by  $\frac{1}{273}$  of its original volume at  $0^\circ\text{C}$ .

### Scale of Thermometry:

12. What is Fahrenheit Scale? Write formula to convert centigrade scale to Fahrenheit scale?

Ans: Fahrenheit is a scale of thermometry which is used to measure temperature. The melting point of ice at 1 atmospheric pressure has a mark  $32^\circ\text{F}$  and that of boiling water is  $212^\circ\text{F}$ . The space between these temperature mark is divided into 180 equal parts and each part is  $1^\circ\text{F}$ .

$$\text{Formula: } ^\circ\text{F} = \frac{9}{5}(^\circ\text{C}) + 32$$

13. Convert  $37^\circ\text{C}$  into  $F^\circ$  scale.

$$\begin{aligned} \text{Ans: } F^\circ &= \frac{9}{5}(C^\circ) + 32 \\ F^\circ &= \frac{9}{5}(37) + 32 \\ F^\circ &= 66.6 + 32 \\ F^\circ &= 98.6 \end{aligned}$$

### General Gas Equation:

14. Derive the units of "R" in general gas equation when the pressure is in atmosphere and volume in  $\text{dm}^3$ ? (2 times)

Ans: Pressure (P) = 1 atm  
Volume (V) =  $22.414\text{ dm}^3$   
T =  $273.15\text{K}$   
n = 1 mole

By putting the value

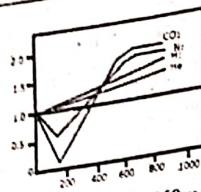
$$R = \frac{PV}{nT} = \frac{1\text{ atm} \times 22.414\text{ dm}^3}{1\text{ mole} \times 273.15\text{K}}$$

$$R = 0.0821\text{ atm} \cdot \text{dm}^3 \cdot \text{K}^{-1} \cdot \text{mole}^{-1}$$

15. Gases show non-ideal behavior at low temperature and high pressure. Justify the statement? (4 times)

Ans: There are considerable intermolecular forces at these conditions (low temperature and high pressure). Thus gases show non-ideal behaviour. At these conditions  $PV = nRT$  does not hold.





16. Calculate density of  $\text{CH}_4$  gas at  $0^\circ\text{C}$  and 760 mm Hg pressure. Molecular mass of  $\text{CH}_4 = 16$ ?

Ans: Density of  $\text{CH}_4$  ( $d$ ) = ?  
 Temperature ( $T$ ) =  $273.16\text{ K}$   
 Pressure ( $P$ ) =  $1\text{ atm}$   
 Molar mass of  $\text{CH}_4$  ( $M$ ) =  $16\text{ g/mole}$   
 $R = 0.0821\text{ dm}^3\text{ atm K}^{-1}\text{ mol}^{-1}$   
 Formula  $d = \frac{PM}{RT}$

By putting values

$$d = \frac{1\text{ atm} \times 16\text{ g/mole}}{0.0821\text{ dm}^3\text{ atm} \cdot \text{K}^{-1}\text{ mol}^{-1} \times 273.16\text{ K}} = 0.71\text{ gdm}^{-3}$$

(4 times)

17. Prove that  $d = \frac{PM}{RT}$   
 Ans: The general gas equation  $PV = nRT$  ..... (i)

Since  $n = \frac{m}{M}$  ..... (ii)

By putting the value on "n" in equation (i)

$$PV = \frac{m}{M} RT$$

$$PM = \frac{m}{V} RT \quad \text{Where } d = \frac{m}{V}$$

So  $PM = dRT$

$$d = \frac{PM}{RT} \quad \text{..... (iii)}$$

This equation can be used to calculate the density of gases.

18. Calculate the value of gas constant  $R$  in S.I units? (21 times)

Ans: Using S.I units of pressure, volume and temperature in general gas equation value of "R" is calculated as:

In S.I units

$$P = 101325\text{ Nm}^{-2}$$

$$V = 0.022414\text{ m}^3$$

By putting the value

$$R = \frac{PV}{nT} = \frac{101325\text{ Nm}^{-2} \times 0.022414\text{ m}^3}{1\text{ mole} \times 273.15\text{ K}}$$

$$R = 8.3143\text{ NmK}^{-1}\text{ mole}^{-1}$$

19. Derive an expression to calculate density of a gas from ideal gas equation. (3 times)

Ans: For calculating the density of ideal gas, we substitute the value of number of moles ( $n$ ) of the gas of the mass ( $m$ ) and the molar mass ( $M$ ) of the gas.

$$PV = nRT$$

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

$$PV = \frac{m}{M} RT$$

Rearranging equation:

$$PM = \frac{m}{V} RT$$

$$PM = dRT \quad (d = \frac{m}{V})$$

$$d = \frac{PM}{RT}$$

20. Derive molecular mass of a gas by general gas equation. (2 times 2018)

Ans: For calculating the molecular mass of ideal gas, we substitute the value of number of moles ( $n$ ) of the gas of the mass ( $m$ ) of the gas.

$$PV = nRT$$

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

$$PV = \frac{m}{M} RT$$

Rearranging equation:

$$M = \left(\frac{m}{V}\right) \frac{RT}{P}$$

$$M = d \frac{RT}{P} \quad (d = \frac{m}{V})$$

**Avogadro's Law:**

21. Define Avogadro's law of gases? (5 times)

Ans: Avogadro's law: "Equal volume of all the gasses at same temperature and pressure contain equal number of molecules"

$$V \propto n$$

For example: 1 mole of  $\text{CO}_2 = 44\text{ g}$  of  $\text{CO}_2 = 6.02 \times 10^{23}$  particles =  $22.414\text{ dm}^3$  at S.T.P.

**Dalton's Law and its Applications:**

22. Explain procedure of sea divers breath? (3 times)

Ans: Deep sea divers cannot breathe with normal air in their tanks. Instead they use a mixture of an inert gas and  $\text{O}_2$  in which partial pressure of oxygen is adjusted around the required limit. Actually in sea after 100 feet depth the diver experiences 3 atm pressures. Hence normal air cannot breathe.

23. Pilots feel uncomfortable breathing at high altitude. Justify? (3 times)

Ans: At higher altitude the partial pressure of oxygen is low. It makes breathing difficult. That's why pilots feel uncomfortable.

24. Define Dalton's law of partial pressures. Give mathematical expression? (2 times)

Ans: Dalton law of partial pressure state that the total pressure exerted by a mixture of non-reacting gasses is equal to the sum of the individual partial pressure of all the

gasses present in a mixture. For example three gases has partial pressure  $P_1$ ,  $P_2$  and  $P_3$ .

So according to Dalton's law

$$P_{\text{Total}} = P_1 + P_2 + P_3$$

25. Derive an expression to find out the partial pressure of a gas?

Ans: In a mixture of gasses partial pressure of any gas can be calculated. Consider two gasses A and B. Let total pressure of mixture is  $P_t$  and number of moles  $n_t$ . While

partial pressure of gas A is  $P_A$  and partial pressure of gas B is  $P_B$ . Then we can

write:  $P_t V = n_t RT$  ..... (i)

$$P_A V = n_A RT \quad \text{..... (ii)}$$

$$P_B V = n_B RT \quad \text{..... (iii)}$$

Divide equation (ii) by equation (i)

$$\frac{P_A V}{P_t V} = \frac{n_A RT}{n_t RT} \quad \frac{P_A}{P_t} = \frac{n_A}{n_t}$$

$$P_A = \frac{n_A}{n_t} P_t$$

OR  $P_A = X_A P_t$

Similarly for gas "B"

$$P_B = X_B P_t$$

26. How Dalton's law of partial pressure is useful in determining pressure of a gas collected over water? (6 times)



30. Total pressure exerted by the mixture of non-reacting gasses is equal to sum of partial pressure of individual gas.

$$P_t = P_1 + P_2 + P_3 \dots$$

Dalton's law of partial pressure is helpful to determine the partial pressure of water vapours or aqueous tension.

$$P_t = P_{H_2O} + P_{gr}$$

$$P_{H_2O} = P_t - P_{gr}$$

27. State Dalton's law of partial pressure. Write its two applications? (2 times)

Ans: Dalton law of partial pressure state that the total pressure exerted by a mixture of non-reacting gasses is equal to the sum of the partial pressure of all the gasses present in a mixture.

- At higher altitudes the pilots feel uncomfortable breathing because the partial pressure of oxygen in the un-pressurized cabin is low as compared to 159 torr, where one feels comfortable breathing.
- Deep sea divers take oxygen mixed with an inert gas and adjust the partial pressure of oxygen according to requirement.

28. Why do we feel comfortable in expressing the densities of gases in unites of  $g\ dm^{-3}$  rather than  $g\ cm^{-3}$ , a unit which is used to express the densities of liquids and solids?

Ans: We feel comfortable in expressing the densities of gases in unites of  $g\ dm^{-3}$  rather than  $g\ cm^{-3}$  because gases have low masses and more volume is required to occupy. Gases are the lightest form of matter. Their densities are very low about 1000 times less than that of liquids and solids.

### Kinetic Molecular Theory:

29. List four postulates of kinetic theory of gases? (5 times) (Home work)

Ans: i) All gasses consists of very small particles called molecules. Gasses like He, Ne, Ar have monoatomic molecules.  
ii) The molecules of gasses are widely separated from one another therefore large empty space exist between them.  
iii) The actual volume of gas molecules is negligible as compared to total volume of the gas.  
iv) Gas molecules are in constant random motion.

30. Describe two causes of deviation from ideality? (8 times)

Ans: Cause of deviation:

- There are no attractive forces among the gas molecules.
- Actual volume of gas molecules is negligible as compared to total volume of the vessel.

31. Apply kinetic molecular theory of gases to explain the Avogadro's law? (3 times) (Home work)

Ans: Consider two gasses 1 and 2 having same "P" and "V". Their number of molecules be  $N_1$  and  $N_2$  and masses  $m_1$  and  $m_2$ . There kinetic equation can be written as:

$$\text{For gas 1 } PV = \frac{1}{3} m_1 N_1 c_1^2$$

$$\text{For gas 2 } PV = \frac{1}{3} m_2 N_2 c_2^2$$

$$\text{So } \frac{1}{3} m_1 N_1 c_1^2 = \frac{1}{3} m_2 N_2 c_2^2$$

$$m_1 N_1 c_1^2 = m_2 N_2 c_2^2 \dots\dots\dots (i)$$

When temperature is same than K.E of both gasses will also same so:

$$\frac{1}{2} m_1 N_1 c_1^2 = \frac{1}{2} m_2 N_2 c_2^2$$

$$m_1 N_1 c_1^2 = m_2 N_2 c_2^2 \dots\dots\dots (ii)$$

Divide equation (ii) by equation (i)

$$N_1 = N_2$$

Hence equal volumes of all the gasses at same temperature and pressure contain equal number of molecules.

Derive Graham's law of diffusion in the light of kinetic molecular theory of gases? (2 times) (Home work)

Ans: According to kinetic molecular equation:

$$PV = \frac{1}{3} m N C^2 \dots\dots\dots (i)$$

If we take one mole of a gas having Avogadro's number of molecules  $N_A$  then the equation (i) can be written as:

$$PV = \frac{1}{3} m N_A C^2 \quad (M = m N_A)$$

M = molar mass of gas

$$C^2 = \frac{3PV}{M} \dots\dots\dots (ii)$$

We know that  $d = \frac{M}{V}$  and  $\frac{1}{d} = \frac{V}{M}$

Putting the value of  $\frac{V}{M}$  in equation (ii)

$$C^2 = \frac{3P}{d}$$

$$\sqrt{C^2} = \sqrt{\frac{3P}{d}}$$

Where  $\sqrt{C^2} \propto r$

$$C_{rms} = \sqrt{\frac{3P}{d}}$$

$$r \propto \frac{1}{\sqrt{d}}$$

That is called Graham's law.

33. Explain Boyle's law in the light of kinetic molecular theory of gases? (2 times) (Home work)

Ans: According to kinetic equation for ideal gas is:

$$PV = \frac{1}{3} m N C^2 \dots\dots\dots (i)$$

According to kinetic molecular theory of gasses the kinetic energy of gas molecules

i.e  $\frac{1}{2} m N C^2$  is directly proportional to absolute temperature.

$$\frac{1}{2} m N C^2 \propto T$$

$$\frac{1}{2} m N C^2 = KT \dots\dots\dots (ii)$$

Multiplying and divide equation (i) by 2

$$PV = \frac{2}{2} \left( \frac{1}{2} m N C^2 \right)$$

$$PV = \frac{2}{3} \left( \frac{1}{2} m N C^2 \right) \dots\dots\dots (iii)$$



Putting equation (ii) in equation (iii)

$$PV = \frac{2}{3}KT$$

If  $T = \text{constant}$  then  $PV = K$

This is Boyle's law. It shows that at constant "T" volume of gas is inverse of pressure.

34. Give two faulty postulates of Kinetic Molecular Theory.

Ans: (i). There are no forces of attraction among the molecules of a gas.  
(ii). The actual volume of the gas molecules is negligible as compared to volume of the vessel.

35. Derive Charles's law by kinetic equation of gases. (2 times)

$$PV = \frac{2}{3}RT$$

$$V = \frac{2}{3} \frac{RT}{P} = \left(\frac{2R}{3P}\right)T$$

At constant pressure,  $\left(\frac{2R}{3P}\right) = k$

$$\frac{V}{T} = k \quad (\text{which is called Charles's law})$$

36. Give two faulty assumption of kinetic molecular theory of gases? (2 times)

Ans: (i) The actual volume of molecules of a gas is negligible as compared to the volume of the gas.

(ii) The molecules of a gas have no forces of attraction for each other.  
37. Hydrogen and Helium are ideal at room temperature but  $\text{SO}_2$  and  $\text{Cl}_2$  are not ideal. Explain? (3 times)

Ans: Both  $\text{SO}_2$  and  $\text{Cl}_2$  have strong intermolecular forces between their molecules at room temperature due to greater size and greater polarizability. Factor "b" for these gases is very high. That's why  $\text{SO}_2$  and  $\text{Cl}_2$  are non-ideal.

$\text{H}_2$  and He both have small size and lesser polarizability. Due to which they have very weak forces of attraction and hence are ideal at room temperature.

**Vander Waal's Equation:**

38. Pressure of  $\text{NH}_3$  gas at given condition is less as calculated by Vander Waal's equation than that of calculated by general gas equation?

Ans: Real gases always show somewhat less pressure as compared to an ideal gas at same conditions of temperature.  $\text{NH}_3$  is a real gas which possesses significant intermolecular forces. The force of attractions reduces the pressure. Hence, observed pressure is lower, which is calculated by Vander Waal's equation.

39. Give units of Vander wall's constants "a" and "b"? (6 times)

Ans: (i) Units of "a"

$$P = \frac{an^2}{V^2}$$

$$a = \frac{PV^2}{n^2} = \frac{\text{atm}(\text{dm}^3)^2}{(\text{mol})^2} = \text{atm} \cdot \text{dm}^6 \cdot \text{mol}^{-2}$$

For S.I units:

$$P = \text{Nm}^{-2} \quad V = \text{m}^3$$

$$a = \frac{PV^2}{n^2} = \frac{\text{Nm}^{-2}(\text{m}^3)^2}{(\text{mol})^2} = \frac{\text{Nm}^{-2}(\text{m}^6)}{(\text{mol})^2} = \frac{\text{Nm}^4}{(\text{mol})^2} = \text{Nm}^4 \cdot \text{mol}^{-2}$$

(ii) Units of "b"

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

$$b = \frac{\text{dm}^3}{\text{mol}} = \text{dm}^3 \cdot \text{mol}^{-1}$$

For S.I units:

$$b = \frac{\text{m}^3}{\text{mol}} = \text{m}^3 \cdot \text{mol}^{-1}$$

### Plasma:

40. Define plasma state. Give its one application? (21 times)

Ans: **Plasma state:** Plasma is a mixture of neutral particles, positive ions and negative electrons. Or

Plasma is a substance in which many of the atoms or molecules are effectively ionized allowing the charge to flow freely.

Applications: (From daily life) (5 times)

i) Plasma can be used for cleaning and sterilization of operation theaters and food.  
ii) Plasma can be used to destroy bacteria, virus, fungi etc.

41. Write name of fourth state of matter. How can it be obtained? (2 times)

Ans: Plasma is the fourth state of matter. It is obtained by heating the substance at high temperature. On heating substances change into atomic state and on further heating it changes to ions.

42. State what is natural and artificial plasma? (2 times)

Ans: **Artificial plasma:** Artificial plasma can be created by using electrical charges on a gas as neon signs.

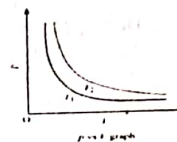
**Natural plasma:** Natural plasma exists only at very high temperature or low temperature vacuum. Natural plasma does not react rapidly.

43. Give two characteristics of plasma. (4 times)

Ans: (i) A plasma must have sufficient number of charged particles so as a whole it exhibits a collective response to electric and magnetic fields. The motion of the particles in the plasma generates fields and electric currents from within plasma density. It refers to the density of the charged particles. This complex set of interactions of electrons and ions are fascinating and complex state of matter.  
(ii) Although plasma includes electrons and ions and conducts electricity, it is macroscopically neutral. In measurable quantities the number of electrons and ions are equal.

44. Why the graph plotted between pressure and volume moves away from pressure axis at higher temperature?

Ans: The reason is that at higher temperature, the volume of the gas has increased. Similarly, if we increase the temperature further, make it constant and plot another isotherm. It further goes away from both axes.



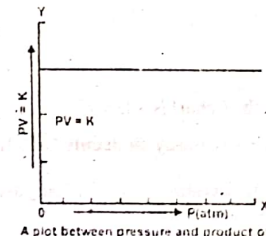
45. Why deep sea divers do not use normal air for breathing in depth of sea?

Ans: Deep sea divers take oxygen mixed with an inert gas say "helium" adjust the partial pressure of oxygen according to the requirement. Actually, in sea after every 100 feet depth, the diver experiences approximately 3 atm pressure, so normal air cannot be breathed in depth of sea. Moreover, the pressure of  $\text{N}_2$  increases in depth of sea and it diffuses in the blood.

46. Why is the plot of PV versus P a straight line at constant temperature and with a fixed number of moles of ideal gas?

Ans: Graph between P and PV: (Proof of  $PV = K$ )

Plot of PV versus P at constant temperature and with fixed number of moles of an ideal gas is a straight line as shown below:



This straight line indicates that  $PV = K$  at all pressure and indicates ideality of gases.



- (vi) Temperature is lowered and pressure is increased.  
 b. Temperature is increased and pressure is lowered.  
 c. Temperature and pressure both are lowered.  
 d. Temperature and pressure both are increased.
- (vii) The molar volume of CO<sub>2</sub> is maximum at  
 a. STP  
 b. 127°C and 1 atm  
 c. 0°C and 2 atm  
 d. 273°C and 2 atm
- (viii) The order of the rate of diffusion of gases NH<sub>3</sub>, SO<sub>2</sub>, Cl<sub>2</sub>, and CO<sub>2</sub> is:  
 a. NH<sub>3</sub> > SO<sub>2</sub> > Cl<sub>2</sub> > CO<sub>2</sub>  
 b. NH<sub>3</sub> > CO<sub>2</sub> > SO<sub>2</sub> > Cl<sub>2</sub>  
 c. Cl<sub>2</sub> > SO<sub>2</sub> > CO<sub>2</sub> > NH<sub>3</sub>  
 d. NH<sub>3</sub> > CO<sub>2</sub> > Cl<sub>2</sub> > SO<sub>2</sub>
- (ix) Equal masses of methane and oxygen are mixed in an empty container at 25°C. The fraction of total pressure exerted by oxygen is  
 a. 1/3  
 b. 8/9  
 c. 1/9  
 d. 16/17

Q2: Fill in the blanks

- (i). The product PV has the S.I. unit of Nm.  
 (ii). Eight grams each of O<sub>2</sub> and H<sub>2</sub> at 27°C will have total K.E in the ratio of 1:16.  
 (iii). Smell of the cooking gas during leakage from a gas cylinder is due to the property of diffusion.  
 (v). The temperature above which a substance exists only as a gas is called critical temperature.

Q3: Label the follow in g sentences as True or False.

- (i). Kinetic energy of molecules of a gas is zero at 0°C. (False)  
 (ii). A gas in a closed container will exert much higher pressure at the bottom due to gravity than at the top. (False)

Q.4 (a) What is Boyle's law of gases? Gives its experimental verification.

Ans: Please consult text book pg # (41-42)

Q.5(b). A sample of carbon monoxide gas occupies 150.0 mL at 25.0°C. It is then cooled at constant pressure until it occupies 100.0 mL. What is the new temperature? (Home work)

Ans:

$$V_1 = 150 \text{ mL} \quad V_2 = 100 \text{ mL}$$

$$T_1 = 25^\circ\text{C} + 273 = 298 \text{ K} \quad T_2 = ?$$

According to Charles's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } T_2 = V_2 \times \frac{T_1}{V_1} = 100 \times \frac{298}{150} = 198.7 \text{ K}$$

(c) Do you think that the volume of any quantity of a gas becomes zero at -273.15°C. Is it not against the law of conservation of mass? How do you deduce the law of absolute zero from this information?

Ans: No volume of a gas cannot be zero at -273°C. Because it is against the law of conservation of mass which states

Mass can neither be created nor destroyed.

Since temperature -273°C is unattainable and it is lowest temperature. Therefore it is taken as Absolute zero of Kelvin scale. Thus, absolute zero is defined as. The hypothetical temperature at which the volumes of all gases become zero called absolute zero.

7. (d) Why do we feel comfortable in expressing the densities of gases in the units of  $\text{g/dm}^3$  rather than  $\text{g/cm}^3$ ? (Home work)

Ans: In gases, molecules are widely separated from each other, and small amount of gas is present in a large volume. Therefore, values of densities will be much smaller expressed in  $\text{cm}^3$ .

Thus, a bigger unit  $\text{dm}^3$  is used because appreciable amount of gas will be present per  $\text{dm}^3$  of gas and values of densities will be greater. E.g. density of CH<sub>4</sub> at STP is  $0.71 \text{ g/dm}^3$  but in  $\text{g/cm}^3$ , it is  $7.1 \times 10^{-4}$  which is inconvenient.

8. Derive the units for general gas constant R in general gas equation. (Class work)

(a) When the pressure is in atmospheres and volume in  $\text{dm}^3$

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

$$\text{Thus } R = \frac{\text{atm} \times \text{dm}^3}{\text{mol} \times \text{K}} = \text{atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

(b) When the pressure is in  $\text{N/m}^2$  and volume in  $\text{m}^3$ .

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

$$\text{Thus } R = \frac{\text{Nm}^{-2} \times \text{m}^3}{\text{mol} \times \text{K}} = \text{Nm mol}^{-1} \text{ K}^{-1} = \text{J mol}^{-1} \text{ K}^{-1}$$

(c) When energy is expressed in ergs.

In SI system, The units of R in terms of energy are  $R = \text{J mol}^{-1} \text{ K}^{-1}$

In CGS system the energy is expressed in units of erg and

$$1 \text{ J} = 10^7 \text{ erg}$$

Therefore, units of R will be  $R = \text{erg mol}^{-1} \text{ K}^{-1}$

However, numerical value of R will be different than SI units.

Q.9(b). Do you think that 1 mole of H<sub>2</sub> and 1 mole of NH<sub>3</sub> at 0°C and 1 atm pressure will have Avogadro's number of particles? (Class work)

Ans: 1 mole of every substance has Avogadro's number of particles. Thus 1 mole of each of H<sub>2</sub> and N<sub>2</sub> will also have Avogadro's number of particles i.e.  $6.02 \times 10^{23}$  particles.

C. Justify that 1 cm<sup>3</sup> of H<sub>2</sub> and 1 cm<sup>3</sup> of CH<sub>4</sub> at STP will have same number of molecules, when one molecule of CH<sub>4</sub> is 8 times heavier than that of hydrogen. In gases distance between two molecule is approximately 300 times its molecular size. Thus volume occupied by gas molecules does not depend upon the molecular size or mass of molecules. The volume of gas only depends upon the number of molecules. Hence according to Avogadro's law equal volume of H<sub>2</sub> and CH<sub>4</sub> at STP will have same number of molecules, although CH<sub>4</sub> molecule is 8 times heavier than H<sub>2</sub> molecule.

Q10. (a). Dalton's law of partial pressures is only obeyed by those gases which don't have attractive forces among their molecules. Explain it. (Home work)

Ans: Dalton's law is an ideal gas law which assumes that there are no attractive or repulsive forces among the gas molecules. Hence every gas molecule moves independently.

On mixing different gases, if they have negligible attractions for each other, then every gas molecule moves independently, exerts its full pressure and thus obeys Dalton's law.

However, if gases develop strong forces for each other on mixing, then their molecules will not be independent in their motion. Therefore, pressure exerted by each gas will be different than expected. Hence, gases with strong attractive forces do not obey Dalton's law.

Q16 Helium gas in a 100 cm<sup>3</sup> container at a pressure of 500 torr is transferred to a container with a volume of 250 cm<sup>3</sup>. What will be the new pressure. (Class work)

a. If no change in temperature occurs

Sol: When the T is not changed.

$$\begin{array}{lll} V_1 & = & 100 \text{ cm}^3 \\ P_1 & = & 500 \text{ torr} \end{array} \quad \begin{array}{lll} V_2 & = & 250 \text{ cm}^3 \\ P_2 & = & ? \end{array}$$

According to Boyle's law

$$P_1 V_1 = P_2 V_2$$

$$P_2 = \frac{P_1 V_1}{V_2} = \frac{500 \times 100}{250} = 200 \text{ torr}$$

b. If its temperature changes from 20°C to 15°C?



Sol:  $V_1 = 100\text{cm}^3$   $V_2 = 250\text{cm}^3$   
 $P_1 = 500\text{ torr}$   $P_2 = ?$   
 $T_1 = 20^\circ\text{C} + 273 = 293\text{K}$   $T_2 = 15^\circ\text{C} + 273 = 288\text{K}$

According to general gas equation.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_2 = \frac{P_1 V_1}{T_1} \times \frac{T_2}{V_2} = \frac{500 \times 100}{293} \times \frac{288}{250} = 196.6\text{ torr}$$

Q.7 a. What are the densities in  $\text{kg/dm}^3$  of the following gases at STP ( $P = 101325\text{ Nm}^{-2}$ ;  $T = 273\text{ K}$ , molecular masses are in  $\text{kg mol}^{-1}$ )

(i) methane, (ii) oxygen, (iii) hydrogen (Class work)

Sol: (i) Methane  $P = 101325\text{ Nm}^{-2}$   
 $T = 273\text{K}$   
Molar mass  $= M = 16\text{g mol}^{-1}$

$$d = \frac{16}{1000} \text{ kg mol}^{-1} = 0.016 \text{ kg mol}^{-1}$$

$$R = 8.314 \text{ J mol}^{-1}\text{K}^{-1} \quad d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.016}{8.314 \times 273} = 0.714 \text{ kgm}^{-3}$$

(ii) Oxygen  $P = 101325\text{ Nm}^{-2}$   
 $T = 273\text{K}$   
Molar mass  $= M = 32\text{g mol}^{-1}$

$$d = \frac{32}{1000} \text{ kg mol}^{-1} = 0.032 \text{ kg mol}^{-1}$$

$$R = 8.314 \text{ J mol}^{-1}\text{K}^{-1} \quad d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.032}{8.314 \times 273} = 1.428 \text{ kgm}^{-3}$$

(iii) Hydrogen  $P = 101325\text{ Nm}^{-2}$   
 $T = 273\text{K}$   
Molar mass  $= M = 2\text{g mol}^{-1}$

$$d = \frac{2}{1000} \text{ kg mol}^{-1} = 0.002 \text{ kg mol}^{-1}$$

$$R = 8.314 \text{ J mol}^{-1}\text{K}^{-1} \quad d = ?$$

The density of gas is given by

$$d = \frac{PM}{RT} = \frac{101325 \times 0.002}{8.314 \times 273} = 0.0893 \text{ kgm}^{-3}$$

b. Compare the values of densities in proportion to their molar masses.

Sol: Density of a gas is directly proportional to its molar masses.

Since molar masses of  $\text{O}_2$ ,  $\text{CH}_4$  and  $\text{H}_2$  are in the order.

$\text{O}_2 > \text{CH}_4 > \text{H}_2$

Hence under given conditions, their densities are also in the same order.

c. How do you justify that increase of volume upto  $100\text{ dm}^3$  at  $27^\circ\text{C}$  of 2 moles of  $\text{NH}_3$  will allow the gas behave ideally, as compared to S.T.P conditions.

Sol: By increasing the volume of  $\text{NH}_3$ . The molecules of  $\text{NH}_3$  become widely separated. As a result forces of attraction between  $\text{NH}_3$  molecules become less and hence it behaves ideally.

Q18 A sample of krypton with a volume of  $6.25\text{ dm}^3$ , a pressure of  $765\text{ torr}$  and a temperature of  $20^\circ\text{C}$  is expanded to a volume of  $9.55\text{ dm}^3$  and a pressure of  $375\text{ torr}$ . What will be its final temperature in  $^\circ\text{C}$ ? (Class work)

Sol:  $V_1 = 6.25\text{ dm}^3$   $V_2 = 9.55\text{ dm}^3$   
 $P_1 = 765\text{ torr}$   $P_2 = 375\text{ torr}$   
 $T_1 = 20^\circ\text{C} + 273 = 293\text{K}$   $T_2 = ?$

According to general gas equation.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$T_2 = \frac{T_1}{P_1 V_1} \times P_2 V_2 = \frac{293}{765 \times 6.25} \times 375 \times 9.55 = 219.46\text{ K}$$

$$\text{Hence } T \text{ in } ^\circ\text{C} = 219.46 - 273 = -53.7^\circ\text{C}$$

Q19 Working at a vacuum line, a chemist isolated a gas in a weighing bulb with a volume of  $255\text{ cm}^3$ , at a temperature of  $25^\circ\text{C}$  and under a pressure of  $10.0\text{ torr}$ . The gas weighed  $12.1\text{ mg}$ . What is the molecular mass of this gas? (Class work)

Sol:  $V = 255\text{ cm}^3 = \frac{255}{1000} \text{ dm}^3 = 0.255\text{ dm}^3$   
 $T = 25^\circ\text{C} + 273 = 298\text{K}$   
 $P = 10\text{ torr} = \frac{10}{760} \text{ atm} = 0.01316\text{ atm}$

$$\text{Mass of gas} = m = 12.1\text{ mg} = \frac{12.1}{1000} \text{ g} = 0.0121\text{ g}$$

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1}\text{K}^{-1}$$

$$\text{Mol. Mass of gas} = M = ?$$

According to general gas equation

$$M = \frac{mRT}{PV} = \frac{0.0121 \times 0.0821 \times 298}{0.01316 \times 0.255} = 88.23 \text{ g/mol}$$

Q20 What pressure is exerted by a mixture of  $2.00\text{g}$  of  $\text{H}_2$  and  $8.00\text{g}$  of  $\text{N}_2$  at  $273\text{K}$  in a  $10\text{ dm}^3$  vessel? (Class work)

Sol: Mass of  $\text{H}_2 = 2\text{g}$   
Moles of  $\text{H}_2 = n_{\text{H}_2} = \frac{2}{2} = 1\text{ mole}$   
Mass of  $\text{N}_2 = 8\text{g}$   
Moles of  $\text{N}_2 = n_{\text{N}_2} = \frac{8}{28} = 0.286\text{ mole}$   
Total number of moles  $= n_t = n_{\text{H}_2} + n_{\text{N}_2} = 1 + 0.286 = 1.286\text{ moles}$

$$\text{Volume of mixture} = V = 10\text{ dm}^3$$

$$T = 273\text{K}$$

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1}\text{K}^{-1}$$

$$\text{Total pressure } P_t = ?$$

According to General Gas Equation.

$$P_t V = n_t RT$$

$$P_t = \frac{n_t RT}{V} = \frac{1.286 \times 0.0821 \times 273}{10} = 2.88\text{ atm}$$

Q22 Calculate the number of molecules and the number of atoms in the given amounts of each gas (Class work)

(a)  $20\text{ cm}^3$  of  $\text{CH}_4$  at  $0^\circ\text{C}$  and pressure of  $700\text{ mm}$  of mercury



$$\frac{20}{1000} \text{ dm}^3 = 0.02 \text{ dm}^3$$

V<sub>1</sub>P<sub>1</sub>

T

$$T = 273 \text{ K}$$

$$P_1 = \frac{700}{760} \text{ atm} = 0.921 \text{ atm}$$

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

According to general gas equation.

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{0.921 \times 0.02}{0.0821 \times 273} = 0.000822 \text{ mol}$$

1 mole of CH<sub>4</sub> contains molecules0.000822 mole of CH<sub>4</sub> contains molecules1 molecule of CH<sub>4</sub> contains atoms4.95 × 10<sup>25</sup> molecule of CH<sub>4</sub> contains atoms

$$= 6.02 \times 10^{23}$$

$$= 6.02 \times 10^{23} \times 0.000822$$

$$= 4.95 \times 10^{25} \text{ molecules}$$

$$= 5$$

$$= 5 \times 4.95 \times 10^{20}$$

$$= 2.475 \times 10^{21} \text{ atoms}$$

(b) 1 cm<sup>3</sup> of NH<sub>3</sub> at 100 °C and pressure of 1.5 atm

$$\text{Sol: } V = 1 \text{ cm}^3 = \frac{1}{1000} \text{ dm}^3 = 0.001 \text{ dm}^3$$

$$T = 100^\circ\text{C} + 273 = 373 \text{ K}$$

$$P = 1.5 \text{ atm}$$

$$R = 0.0821 \text{ atm dm}^3 \text{ mol}^{-1} \text{ K}^{-1}$$

n = ?

According to general gas equation.

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{1.5 \times 0.001}{0.0821 \times 373} = 4.9 \times 10^{-5} \text{ mol}$$

1 mole of NH<sub>3</sub> contains molecules0.000049 mole of NH<sub>3</sub> contains molecules1 molecule of NH<sub>3</sub> contains atoms2.95 × 10<sup>19</sup> molecule of NH<sub>3</sub> contains atoms

$$= 6.02 \times 10^{23}$$

$$= 6.02 \times 10^{23} \times 0.000049$$

$$= 2.95 \times 10^{19} \text{ molecules}$$

$$= 4$$

$$= 4 \times 2.95 \times 10^{19}$$

$$= 1.18 \times 10^{20} \text{ atoms}$$

Q23 Calculate the masses of 10<sup>20</sup> molecules of each of H<sub>2</sub>, O<sub>2</sub>, and CO, at STP. What will happen to the masses of these gases, when the temperature of these gases are increased by 100 °C and the pressure is decreased by 100 torr. (Class work)Sol: For H<sub>2</sub>Number of molecules = 10<sup>20</sup>

Mass of hydrogen = ?

Since 6.02 × 10<sup>23</sup> molecules of H<sub>2</sub> = 2g

$$\text{Therefore } 10^{20} \text{ molecules of H}_2 = \frac{2}{6.02 \times 10^{23}} \times 10^{20} = 3.322 \times 10^{-4} \text{ g}$$

Number of molecules = 10<sup>20</sup>

Mass of Oxygen = ?

Since 6.02 × 10<sup>23</sup> molecules of O<sub>2</sub> = 32g

$$\text{Therefore } 10^{20} \text{ molecules of O}_2 = \frac{32}{6.02 \times 10^{23}} \times 10^{20} = 5.3152 \times 10^{-3} \text{ g}$$

Number of molecules = 10<sup>20</sup>

Mass of Carbon dioxide = ?

Since 6.02 × 10<sup>23</sup> molecules of CO<sub>2</sub> = 44g

$$\text{Therefore } 10^{20} \text{ molecules of CO}_2 = \frac{44}{6.02 \times 10^{23}} \times 10^{20} = 7.3084 \times 10^{-3} \text{ g}$$

## CHAPTER-4 (LIQUIDS AND SOLIDS)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS-2020)

### Liquids

Dipole – dipole forces are present among:

1. (a) Molecules of Iodine (b) Atoms of Neon in gaseous state  
(c) Chloroform molecules (d) CCl<sub>4</sub> molecules

2. Liquid hydrocarbon is:

- (a) Methane (b) Propane (c) Ethane (d) Hexane

3. Water has maximum density at:

- (a) 4.0°C (b) 0°C (c) 100°C (d) 10°C

4. Which of the given has Hydrogen Bonding:

- (a) CH<sub>4</sub> (b) CCl<sub>4</sub> (c) NH<sub>3</sub> (d) NaCl

5. Hydrogen bonding is maximum in:

- (a) HI (b) HBr (c) HCl (d) HF

6. Which of the following liquid has highest boiling point:

- (a) HCl (b) HBr (c) H<sub>2</sub>O (d) Br<sub>2</sub>

7. Ice occupies more spaces than liquid water upto:

- (a) 9% (b) 10% (c) 11% (d) 12%

8. When water freeze its volume increases:

- (a) 10% (b) 9% (c) 15% (d) 18%

### Solids

9. Allotropy is the property of

- (a) element (b) compound (c) Mixture (d) Ions

10. Existence of an element in more than one crystalline form is called: (1 time)

- (a) Allotropy (b) Isotropy (c) Isomorphism (d) Polymorphism

11. Transition temperature of KNO<sub>3</sub> is:

- (a) 13.2°C (b) 95.5°C (c) 128°C (d) 32.2°C

12. Transition temperature of Tin(Sn) is:

- (a) 128°C (b) 95.5°C (c) 13.2°C (d) 32.8°C

13. The number of Cl<sup>-</sup> ions per unit cell of NaCl are:

- (a) 8 (b) 6 (c) 4 (d) 2

14. Shape of Graphite is:

- (a) Hexagonal (b) Cubic (c) Tetragonal (d) Monoclinic

15. Diamond and graphite are example of:

- (a) Isomorphism (b) Polymorphism (c) Isomerism (d) Allotropy

16. In Al<sub>2</sub>O<sub>3</sub> the ratio between the ions is:

- (a) 1 : 2 (b) 2 : 1 (c) 2 : 3 (d) 3 : 2

17. The example of hexagonal system is:

- (a) Sulphur (b) NaCl (c) Graphite (d) Diamond

18. When a=b=c and α=β=γ=90° then the crystal system is: (2 times)

- (a) Hexagonal (b) Monoclinic (c) Cubic (d) Tetragonal

19. The Axes(unit cell length) for Cu is:

- (a) a ≠ b = c (b) a ≠ b ≠ c (c) a=b≠c (d) a=b=c

20. Which one of the following is in liquid state at room temperature?

- (A) Methane (B) Ethane (C) Hexane (D) Propane

21. A ring has 6.0 g of diamond in it. Calculate the number of atoms of Carbon in it:

- (A) 6.02 × 10<sup>23</sup> (B) 3.01 × 10<sup>23</sup> (C) 9.03 × 10<sup>23</sup> (D) 1.8 × 10<sup>24</sup>

22. Which impurity makes the shape of sodium chloride crystal needle like:

- (A) MgSO<sub>4</sub> (B) Urea (C) Glucose (D) MgCO<sub>3</sub>

23. The strongest acid among Halogen acids is:

- (A) HCl (B) HBr (C) HI (D) HF

24. Dipole-induced dipole forces are also called:

- (A) Dipole-dipole forces (B) Ion-dipole forces (C) Debye forces (D) London dispersion forces

25. The boiling point of water at Murree hill is:



- (A) 99.8°C (B) 98°C (C) 100°C (D) 89°C.
- 26- Geometry of diamond is: (A) Tetragonal (B) Cubic (C) Rhombohedral (D) none of these
- 27- Density of ice is minimum at 4°C due to: (a) Empty spaces in structure of ice (b) Tetrahedral shape of crystal of ice (c) Large bond lengths (d) Large bond angles
- 28- The solid which has no definite crystalline shape: (a) Sugar (b) Salt (c) Glass (d) Dry ice
- 29- Hydrogen bonding is maximum in: (a) HI (b) HBr (c) HCl (d) H<sub>2</sub>O
- 30- Transition temperature of S<sub>8</sub> (monoclinic)  $\rightleftharpoons$  S<sub>8</sub> (Rhombic) is: (a) 13.2°C (b) 95.5°C (c) 128°C (d) 110°C

## Answers

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
c	d	A	C	d	c	a	B	a	a	c	c	c	a	d
16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
c	c	c	d	c	B	B	c	c	b	b	a	C	d	b

## CHAPTER-4 (LIQUIDS AND SOLIDS)

### SHORT QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

#### Liquids

1. **Intermolecular Forces:**  
 1. Intermolecular forces are weaker than intermolecular forces why? (2 times)  
 Ans: The force present among the atoms of a molecule is called intermolecular forces. These are true chemical bonds while the forces of attraction present among particle of a substance that close them together. These are weak attractions as compare with intermolecular forces.
2. Define dipole-dipole forces with example? (4 times)  
 Ans: The positive end of the molecule attracts the negative end of the other molecule and these electrostatic forces of attraction are called dipole - dipole forces. The examples of molecules which show dipole-dipole attractions are numerous. Two of these are HCl and CHCl<sub>3</sub>.
- $$H^{\delta+} - Cl^{\delta-} \dots\dots H^{\delta+} - Cl^{\delta-} \dots\dots H^{\delta+} - Cl^{\delta-}$$
- $$Cl_3C^{\delta+} - H^{\delta-} \dots\dots Cl_3C^{\delta+} - H^{\delta-} \dots\dots Cl_3C^{\delta+} - H^{\delta-}$$
3. Write down two applications of hydrogen bonding?  
 Ans: (i) Large protein molecules in living organisms are stabilized due to H-bonding.  
 (ii) Double helix of DNA molecule is linked together through H-bond.  
 (iii) Helical structure of proteins is stable due to hydrogen bonding.
4. Why HF is weaker acid than HCl? (4 times)  
 Ans: Strength of an acid depends upon the ionization of an acid. Ionization of HF is low because H<sup>+</sup> is entrapped between two fluorine atoms. From one side it is covalently bonded while from other side strong hydrogen bond is present. That's why HF is a weaker acid than HCl.
5. Ice floats on H<sub>2</sub>O. Give reason? (6 times)  
 Ans: When the temperature of liquid water decreases the arrangement of molecules also changes. In case of ice hexagonal arrangements of H<sub>2</sub>O molecules are formed with large empty spaces than liquid. About 9 % expansions in volume take place. Due to lesser density of ice, it floats over water.
6. Water is liquid at room temperature while H<sub>2</sub>S is a gas. Comment? (3 times)  
 Ans: In H<sub>2</sub>O strong H-bonding is present which make it liquid. But H<sub>2</sub>S have weak intermolecular force. Due to this H<sub>2</sub>S is a gas at room temperature.
7. Describe the importance of vacuum distillation? (5 times)

- Ans: Some compounds decompose at their boiling point. Such compounds are distilled under reduced pressure. In this process, a liquid is made to boil at lower temperature by decreasing pressure, e.g. glycerine.
8. Density of ice is less than liquid water. Explain with reason? (2 times)  
 Ans: In liquid water the molecules are extensively associated with each other due to strong H-bonding. But this association is irregular, when temperature dropped to 0°C. The molecules of water arrange themselves in a regular pattern. Due to this, empty spaces among H<sub>2</sub>O molecules increases and hence density of ice decreases.
9. What are London Dispersion forces? (2 times)  
 Ans: The momentary forces of attraction created between instantaneous dipole and the induced dipole is called dipole- induced dipole interaction or London force. It is very short lived attraction because the electrons keep moving. This movement of electrons cause the dipole to vanish as quickly as they are formed.
10. What is the role of Hydrogen Bonding in Biological Compounds?  
 Ans: Hydrogen bonding exists in the molecules of living system. In proteins like hair, silk and muscles consists of long chains of amino acids due to hydrogen bonding. These long chains are coiled about one another into a spiral. This spiral is called a helix. Such a helix may either be right handed or left handed. DNA has two spiral chains due to hydrogen bonding these are linked together.
11. Define hydrogen bonding and give one example.  
 Ans: Hydrogen bonding is the force of attraction between a highly electronegative atom and partial positively charged hydrogen atom. The hydrogen bonding present in the molecules of ammonia and those of hydrofluoric acid is the example of hydrogen bonding.
12. Describe cleaning action of soaps and detergents on the basis of H-bonding.  
 Ans: Soap and detergents perform the cleansing action because the polar part of their molecules are water soluble due to hydrogen-bonding and the non-polar parts remain outside water, because they are alkyl or benzyl portion and are insoluble in water.
13. What type of intermolecular forces will dominate in the following liquid?  
 (a) CH<sub>3</sub>COCH<sub>3</sub> (propanone) (b) C<sub>8</sub>H<sub>18</sub> (Octane)  
 Ans: (a) CH<sub>3</sub>COCH<sub>3</sub> (propanone) : Dipole-dipole forces.  
 (b) C<sub>8</sub>H<sub>18</sub> (Octane): Instantaneous dipole-induced dipole forces or London dispersion forces.
4. **Liquid Crystals:**  
 14. How liquid crystals act as temperature sensors? (2 times)  
 Ans: Liquid crystals can diffract light, when one of the wavelengths of white light is reflected from a liquid crystal it appears colored. As the temperature changes the distance between the layers of the molecules of liquid crystals changes. Therefore the colour of the reflected light changes. Thus liquid crystal can be used as temperature sensors.
15. Give two important uses of liquid crystals? (12 times)  
 Ans: (i) These are used as temperature sensors.  
 (ii) Liquid crystals are used to find the point of potential failure in electrical circuits.  
 (iii) In chromatographic separations these are used as solvent.
16. Define liquid crystal with one example. (4 times)  
 Ans: A liquid crystalline state exist between two temperatures. i.e. melting temperature and clearing temperature. Those substances which exhibit this liquid crystalline state are called liquid crystals. For example cholesteryl benzoate. This compound turns milky liquid at 145°C and becomes clear liquid at 179°C.

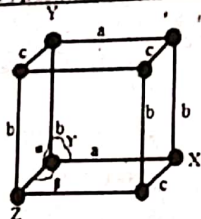
#### Solids

1. **Types and Properties of Crystalline Solids:**  
 17. Define isomorphism with two examples? (7 times)  
 Ans: Isomorphism: The phenomenon in which two different substances exist in the same crystalline form is called isomorphism.  
 e.g. NaNO<sub>3</sub>, KNO<sub>3</sub> Rhombohedral.  
 K<sub>2</sub>SO<sub>4</sub>, K<sub>2</sub>CrO<sub>4</sub> Orthorhombic



17. Define polymorphism with example? (10 times)  
 Ans: Polymorphism: The phenomenon in which a substance exists in more than one crystalline form is called polymorphism.  
 e.g.  $\text{AgNO}_3$  exists in Rhombohedral & Orthorhombic forms.
18. Define crystalline solid and crystallites.  
 Ans: Crystalline solids: Those solids in which atoms, ions or molecules are arranged in a definite three dimensional pattern are called crystalline solid. This recurring regular geometrical pattern of structure extends three dimensionally. Crystallites: The crystalline part of otherwise amorphous solids are called crystallites.
19. What are crystallographic elements? Explain with diagram.  
 Ans: Crystallographic elements: In a unit cell the length and angles of a unit cell are called crystallographic elements. There are six crystallographic elements.

### 1. Types and Properties of Crystalline Solids:



'a' edge length along x-axis  
 'b' edge length along y-axis  
 'c' edge length along z-axis  
 $\alpha$  angle between 'b' and 'c'  
 $\beta$  angle between 'a' and 'c'  
 $\gamma$  angle between 'a' and 'b'

20. Describe crystallographic elements? (2 times)  
 Ans: Crystallographic elements: In a unit cell the length and angles of a unit cell are called crystallographic elements. There are six crystallographic elements. Lengths "a, b and c" and angles " $\alpha$ ,  $\beta$  and  $\gamma$ ".
21. Explain cleavage of the crystals and cleavage plane? (1 time) (2017= 1 time)  
 Ans: Cleavage crystals: Cleavage is a directional property because division of a crystal from one direction may be easy as compared to other. This is due to different arrangement of particles.  
 Cleavage planes: When forces are applied on a crystalline solid it breaks along definite planes called cleavage planes.
22. Transition temperature is the term used for elements as well as compounds. Explain? (2 times)  
 Ans: Transition temperature is a temperature at which two crystalline forms of the same substance can co-exist in equilibrium with each other. At this temperature, one crystalline form (elemental or compound) of a substance changes to another. For example:  
 Grey tin (cubic)  $\rightleftharpoons$  White tin (Tetragonal) Transition temperature =  $13.2^\circ\text{C}$   
 $\text{KNO}_3$  (orthorhombic)  $\rightleftharpoons$   $\text{KNO}_3$  (rhombohedral) Transition temperature =  $128^\circ\text{C}$
23. Define molar heat of fusion and molar heat of vapourization? (2 times)  
 Ans: Molar heat of fusion is the amount of heat absorbed by one mole of a solid when it melts into liquid form at its melting point. The pressure, during the change is kept one atmosphere.  
 Molar heat of vapourization is the amount of heat absorbed when one mole of a liquid is changed into vapours to its boiling point. The pressure, during the change is kept one atmosphere.
24. Define transition temperature with two examples? (7 times)  
 Ans: The temperature at which two crystalline forms of the same substances can co-exist in equilibrium with each other is called transition temperature.  
 Example:  
 Sulphur  $95.5^\circ\text{C}$  Sulphur  
 (Rhombohedral) (Monoclinic)  
 Grey tin  $13.2^\circ\text{C}$  White tin  
 (Cubic) (Tetragonal)

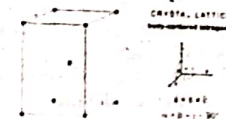
25. Differentiate between an amorphous solid and crystalline solid? (3 times)

	Crystalline solid	Amorphous solid
(i)	In these solids the atoms, ions and molecules are arranged in a definite three dimensional patterns.	In these solids the particles are not arranged in definite three dimensional patterns.
(ii)	These have sharp melting points. e.g. $\text{NaCl}$ , $\text{KNO}_3$	These have not sharp melting points. e.g. Glass, plastic etc.

### 26. Differentiate between isotropy and anisotropy?

	Isotropy	Anisotropy
(i)	The properties which do not depend upon direction are called isotropic, e.g. melting point.	Some of crystals show variation in physical properties depends upon direction. Such properties are called anisotropic properties.

27. Why amorphous solids like glass are also called super cooled liquids? (2 times)  
 Ans: A liquid can exist below its freezing point is said to be undercooled or super-cooled liquid. Liquid glass when cooled, cannot move readily into the position of a regular crystal lattice and they formed amorphous solid with irregular arrangement.
28. Define anisotropy, Give an example? (2 times)  
 Ans: Some of the crystals show variation in physical properties depending upon the direction. Such properties are called anisotropic properties and the phenomenon is referred to as anisotropy. For example the physical properties like refractive index, sometimes anisotropic in nature of some crystals. For example electrical conductivity of graphite is greater in one direction than in another.
29. Explain tetragonal system showing angles, phases and give two examples?  
 Ans: In tetragonal system two axis are of equal length and the third axis is either shorter or larger than the other than two. All angles are  $90^\circ$ .



30. Cleavage of the crystal is itself anisotropic. Give reason? (7 times)  
 Ans: Cleavage of a crystal is a directional property because breakage of a crystal from one direction may be easy as compared to other. This is due to the different arrangement of particles in different direction.  
 For example: Mica crystal can be cleavage parallel to the sheets easily.
31. Define allotropy with an example? (2 times)  
 Ans: Allotropy: The phenomenon in which an element exists in more than one crystalline form is called allotropy.  
 e.g. Carbon (i) Diamond (ii) Graphite
32. Ionic compounds do not show the phenomenon of isomerism. Why?  
 Ans: Cations and anions are arranged in a well-defined geometrical pattern, so they are crystalline solids at room temperature. Ionic crystals are very stable compounds. Ionic crystals don't exist as individual neutral independent molecules, therefore ionic compounds do not show the phenomena of isomerism.
33. Why sodium is softer than copper but both are good conductors of electricity? (2 times)  
 Ans: Both sodium and copper are metals and have free electrons in them. Due to movement of these free electrons, they are good conductor. However, strength of metallic bond depends upon the number of valence electrons. Copper has greater number of valence electrons than sodium, so it is more hard than sodium.

### 3. Units Cell:

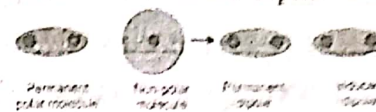
34. Define unit cell with example? (2 times)  
 Ans: Unit cell: The smallest part of a crystal lattice. Showing all the characteristics features of the whole crystal is called unit cell. e.g. Cubic unit cell.
35. Why the metals are malleable and ductile? (2 times)  
 Ans: Malleable means metals can be beaten into sheets. Ductile means metal can be squeezed into wires. When stress is applied on metal, their layers slip passes over



- each other but the strength of metallic bond remain same. Therefore metals are malleable and ductile.
35. **Ionic solids are insulators in solid state, but become conductors when dissolved in water. Explain?** (4 times)  
 Ans: Ionic crystals do not conduct electricity in solid state because ions are held tightly in a crystal lattice. But when dissolve in water then these ions are free to move. Hence ionic crystals conduct electricity in solution form. (9 times)
36. **Ionic crystals are highly brittle. Why?**  
 Ans: Ionic crystals are brittle when some stress is applied to ionic crystals, the layers of ions move and similar ions come in front of each other. These same charges repel to each other and solid break into pieces.
37. **One of the unit cell angles of Hexagonal crystal is  $120^\circ$ ? In hexagonal system two axis are of equal length and are in one plane making an angle of  $120^\circ$ , with each other. The third axis which is different in length than the other two is at right angle to these two axes.**  
 Ans: (4 times)
6. **Lattice Energy:**  
 38. **What is Lattice and Lattice energy?**  
 Ans: Lattice: The location of position of atoms; ions or molecules in solid is called lattice.  
 Lattice energy: The amount of energy required to break one mole of an ionic solid into ions is called lattice energy.  
 e.g.  $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$   $\Delta H^\circ = +787 \text{ KJ/mol}$
7. **Metallic Solids:**  
 39. **How electrical conductivity of the metals decreases by increasing temperature?** (8 times)  
 Ans: When the temperature is increased the positive ions begin to vibrate about their mean positions. These vibrations hinder the movement of free electrons in metals and therefore electrical conductivity is reduced.
40. **Write down any four properties of Ionic solids.**  
 Ans: Properties of Ionic Solids:  
 Properties of ionic solids are as follows:
- Physical State:**  
The cations and anions are arranged in a well-defined geometrical pattern, so they are crystalline solids at room temperature. Under ordinary conditions of temperature and pressure they never exist in the form of liquids or gases.
  - Crystal formation:**  
Ionic solids do not exist as individual neutral independent molecules. Their cations and anions attract each other in all directions and these forces are non-directional. The close packing of the ions enables them to occupy minimum space. A crystal lattice is developed when the ions arrange themselves systematically in a alternate manner.
  - Hard solids with high melting point:**  
Ionic crystals are very stable compounds. Very high energy is required to separate the cations and anions from each other against the forces of attraction. That is why ionic crystals are very hard, have low volatility and high melting and boiling point.
  - Radius ratio and Structure:**  
The structure of the ionic crystals depends upon the radius ratio of cations and anions. For example  $\text{NaCl}$  and  $\text{CsF}$  have the same geometry because the radius ratio in both the cases is the same.
41. **How are liquid crystals used to locate veins, arteries infections and tumors?**  
 Ans: Liquid crystalline substances are used to locate the veins, arteries, infections and tumors. The reason is that these parts of the body are warmer than the surrounding tissues. Specialists can use the techniques of skin thermography to detect blockages in veins and arteries. When a layer of liquid crystal is painted on the surface of the breast, a tumor shows up as a hot area which is coloured blue. This technique has been successful in the early diagnosis of breast cancer.
42. **Lower alcohols are soluble in water but hydrocarbons are insoluble. Give reason.**

Ans: Water is the best example of a hydrogen bonding system. Similarly, ethyl alcohol also has the tendency to form hydrogen bonds. So, ethyl alcohol can dissolve in water because both can form hydrogen bonds with each other. Similarly carboxylic acids are also soluble in water, if their sizes are small. Hydrocarbons are not soluble in water at all, because they are non-polar compounds and there are no chances of hydrogen bonding between water and hydrocarbon molecules.

43.



Ans: Sometimes, we have a mixture of substances containing polar and non-polar molecules. The positive end of the polar molecule attracts the mobile electrons of the nearby non-polar molecule. In this way polarity is induced in non-polar molecule, and both molecules become dipoles. These forces are called dipole-induced dipole forces or as Debye forces. The following figure makes the idea clear.

44.

Ans:



53. Ethane and hexane has Boiling Points,  $-88.6^{\circ}\text{C}$  and  $68.7^{\circ}\text{C}$  respectively. Comment on it.
- Ans: Hexane has greater molecular size. It has molecule with large chain length than ethane. There are more places along its length, where they can be attracted to other molecules. Thus hexane molecules experience stronger London forces. Due to this reason, it has greater boiling point than ethane.

## CHAPTER-4 (LIQUIDS AND SOLIDS)

### LONG QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

#### Liquids

- Define and explain London forces. Describe the factors affecting the London dispersion forces. (1 time)(2017=1 time)
- Explain the term dipole-dipole and dipole induces dipole forces.
- What are intermolecular forces? Write the names of different types of their forces and explain instantaneous dipole-induced dipole forces. (4 times)
- What is H-bonding? Discuss H-bonding in biological compounds. (4 times)
- Define hydrogen bonding. How it is helpful in explaining the structure of ice.
- What is hydrogen bonding? Explain role of hydrogen bonding in food and biological material.
- Explain hydrogen bonding in  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  and  $\text{HF}$ . How it is helpful in explaining structure of ice.
- How does the hydrogen bonding explain the formation of ice lesser density than liquid water? (2 times)
- Define liquid crystals. Discuss important uses of liquid crystal. (6 times)

#### Solids

- Differentiate between isomorphism and polymorphism with suitable example. (2 times)
- Explain seven crystal systems with angles and edges.

### Solved Exercise Chapter # 4 (Liquids)

#### ALP SMART SYLLABUS 2020

- Q1. Choose the best answers from the given choices. (C.W)
- London dispersion forces are the only forces present among the
    - molecules of water in liquid state
    - atoms of helium in gaseous state at high temperature
    - molecules of solid iodine.
    - molecules of hydrogen chloride gas.
  - Acetone and chloroform are soluble in each other due to
    - intermolecular hydrogen bonding
    - ion-dipole interaction
    - instantaneous dipole
    - all of the above
  - $\text{NH}_3$  shows a maximum boiling point among the hydrides of Vth group elements due to
    - very small size of nitrogen
    - lone pair of electrons present on nitrogen.
    - enhanced electronegative character of nitrogen
    - pyramidal structure of  $\text{NH}_3$
  - When water freezes at  $0^{\circ}\text{C}$ , its density decreases due to
    - cubic structure of ice

- (b) empty spaces present in the structure of ice  
(c) change of bond lengths  
(d) change of bond angles
- Q2. Fill in the blanks with suitable words (C.W)
- The polarizability of noble gases increases down the group and results in the increase in their boiling points.
  - Hydrogen bonding is developed in acetone and chloroform when they are mixed together.
  - Exceptionally weak acidic strength of  $\text{HF}$  is due to strong hydrogen bonding present in it.
  - The concept of dynamic equilibrium is the ultimate result of all reversible systems.
  - $\Delta H_v$  of  $\text{C}_6\text{H}_{14}$  should be greater than that of  $\text{C}_2\text{H}_6$ .
  - During the formation of ice from liquid water there is a 9 % increase in volume.
  - The rate of increase of vapour pressure of water increases at high temperatures.
  - A layer of ice on the surface of water insulates the water underneath for further heat loss.
- Q3. Indicate true or false as the case may be (C.W)
- Dipole-dipole forces are weaker than dipole-induced dipole forces. (False)
  - The ion dipole interactions are responsible for the dissolution of an ionic substance in water. (True)
  - The high polarizability of iodine is responsible for its existence in solid form and its difference from other halogens. (True)
  - The strong hydrogen bonding in  $\text{H}_2\text{S}$  makes it different from water. (False)
  - Hydrocarbons are soluble in water because they are polar compounds. (False)
  - The viscosities of liquids partially depend upon the extent of hydrogen bonding. (True)
  - Ice does not show any vapour pressure on its surface at  $-1^{\circ}\text{C}$ . (False)
- Q4 (a) What type of intermolecular forces will dominate in the following liquids.  
(i) Ammonia,  $\text{NH}_3$  (ii) Octane,  $\text{C}_8\text{H}_{18}$  (iii) Argon, Ar (C.W)  
(iv) Propanone,  $\text{CH}_3\text{COCH}_3$  (v) Methanol,  $\text{CH}_3\text{OH}$
- Sol: (i) Ammonia,  $\text{NH}_3$  :-  
N is sufficiently strong E.N. element, therefore, N of one molecule will develop H-bond with H of another molecule of  $\text{NH}_3$ . Hence in  $\text{NH}_3$  H-bonding will dominate.
- (ii) Octane,  $\text{C}_8\text{H}_{18}$  :-  
Octane is a non-polar molecule. It will not have dipole-dipole interaction or H-bonding. Its molecules will have weak London dispersion forces among themselves.
- (iii) Argon, Ar :-  
Argon is noble gas and is non-polar. It will have London dispersion forces among its molecules. Due to non-polar nature. H-bonding or dipole-dipole interactions will not be present.
- (iv) Propanone,  $\text{CH}_3\text{COCH}_3$  :-  
It has a polar carbonyl group  $\text{C}=\text{O}$ . Due to this group, propanone will develop dipole-dipole interaction.
- (v) Methanol,  $\text{CH}_3\text{OH}$  :-  
In methanol, strong electronegative O is present. Thus O atom of one molecule will form H-bond with H of another methanol molecule. Thus, methanol chiefly has H-bonding.
- (b) Propanone ( $\text{CH}_3\text{COCH}_3$ ), propanol ( $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ ) and butane ( $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$ ) have very similar relative molecular masses. List them in the expected order of increasing boiling points. Explain your answer.
- Sol: The b.p will be in the following order:  
Propanol > Propanone > Butane  
 $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$  >  $\text{CH}_3\text{COCH}_3$  >  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$   
 $97^{\circ}\text{C}$  >  $56^{\circ}\text{C}$  >  $0^{\circ}\text{C}$
- It is because  
> Propanol has H-bonding



> Propanone has dipole-dipole interaction  
Butane has weak London dispersion forces  
Since H-bonding is stronger than dipole-dipole interaction, which in turn is stronger than London forces. (C.W)

**Q.5** Explain the following with reasons.  
(i) In the hydrogen bonded structure of HF, which is the stronger bond: the shorter covalent bond or the longer hydrogen bond between different molecules.

**Sol:** Generally shorter bond are stronger bond.  
Further, in HF short covalent bond is formed by overlap of orbitals of H and F. While H-bond is just an electrostatic interaction between different molecules. Hence short covalent bond is stronger than longer H-bond.

(ii) In a very cold winter the fish in garden ponds owe their lives to hydrogen bonding?

**Sol:** In cold winter when temperature falls below 4°C, cold water being lighter comes to the surface. When temperature is further decreased, cold water freezes to ice. Thus an insulating layer of ice is formed above warm water. This layer of ice prevents further heat loss from underneath water. Thus, it save aquatic life e.g. fishes from cold.

(iii) Water and ethanol can mix easily and in all proportions.

**Sol:** Because their molecules can develop hydrogen bonding. (C.W)

**Q6 (a)** Briefly consider some of the effects on our lives if water has only a very weak hydrogen bonding present among its molecules.

**Sol:** In H<sub>2</sub>O if H-bonding is weak, it will have following effects on life.

- The freezing point of water will be lower than 0°C. This will be a serious problem to the life of living things.
- Water acts as thermostat for our earth. With weaker H-bonding, water will not be a thermostat. Thus there will be large changes in the temperature of the earth.
- Due to weak H-bonding density of ice will be less than water. Thus it would be danger for aquatic life.
- With weaker H-bonding, the structures of proteins and DNA would not be stabilize.

(b) All gases have a characteristic critical temperature. Above the critical temperature it is impossible to liquefy a gas. The critical temperatures of carbon dioxide and methane are 31.14 °C and -81.9 °C, respectively. Which gas has the stronger intermolecular forces? Briefly explain your choice?

**Sol:** The critical temperature of CO<sub>2</sub> is closer to room temperature (i.e. 31.1°C). It shows that CO<sub>2</sub> is ready to be liquefied around room temperature by applying pressure. However critical temperature of CH<sub>4</sub> is -81.9°C. It shows that CH<sub>4</sub> must be cooled below this temperature so that it could be liquefied by applying pressure.

Since at room temperature CO<sub>2</sub> is easily liquefied but CH<sub>4</sub> not. Hence, CO<sub>2</sub> has stronger intermolecular forces than CH<sub>4</sub> at the same temperature.

**Q7** Three liquids have the properties mentioned against their names Water (C.W) Propanone Pentane

	Water	Propanone	Pentane
Molecular Formula	H <sub>2</sub> O	C <sub>3</sub> H <sub>6</sub> O	C <sub>5</sub> H <sub>12</sub>
Relative molecular mass	18	58	72
Enthalpy of vaporization	41.1	31.9	27.7
Boiling point (°C)	100	56	36

(a) What type of intermolecular force predominates in each liquid?

(i) water (ii) propanone (iii) pentane

**Sol:** H<sub>2</sub>O has H-bonding.  
Propanone has dipole-dipole interaction.  
Pentane has London dispersion forces.

(b) What do you deduce about the relative strength of these forces in the liquids? Justify your conclusions.

**Sol:** The order of strength of these forces is

H-bonding > dipole-dipole interaction > London dispersion forces.  
Hence, due to strong H-bonding b.p. of H<sub>2</sub>O and its Hv is greater than

propanone, which has dipole-dipole interaction.

In pentane weak London dispersion forces are present, therefore, its boiling point and Hv are lower than both H<sub>2</sub>O and Propanone.

If the liquids are shaken together in pairs,

(c) Which pair would be unlikely to mix?

**Sol:** Water & Pentane will not mix with each other since pentane is non-polar and water is polar. Therefore they will not have attraction for each other.

(ii) Explain this immiscibility in terms of the forces between the molecules.

**Sol:** H<sub>2</sub>O and propanone are polar substances and have H-bonding and dipole-dipole interactions. Hence they attract the molecules of each other. Thus, they mix with each other. However Pentane is non-polar and has weak London dispersion forces. It can not develop forces with water or propanone molecules, hence it is immiscible with H<sub>2</sub>O and propanone.

(iii) Choose one of the pairs that mix and say whether the enthalpy change on mixing would be positive or negative.

**Sol:** Water and propanone will mix with each other. This mixing is exothermic process and have -ve enthalpy change, because both water and acetone have strong attraction for each other during which energy is released.

**Q8** Describe the various forces responsible for keeping the particles together in the following elements and compounds and their effects on physical properties making use of the data below. (C.W)

Substance	Formula	Molar Mass	M.P (°C)
Neon	Ne	20	-248
Argon	Ar	40	-189
Water	H <sub>2</sub> O	18	0
Sodium fluoride	NaF	42	993
Diamond	C	12	3350

**Sol:**

- Ne and Ar are noble gases and have weak London dispersion forces. Since Ar is larger in size, therefore, it has greater polarizability than Ne and therefore strong London forces. Hence its m.p is higher than Ne.
- Water has strong H-bonding, therefore, it will have sufficiently high m.p. than Ne and Ar.
- NaF is an ionic solid. It has cubic crystal structure in which each Na<sup>+</sup> ions. These ions have strong electrostatic forces. This arrangement gives hardness of NaF crystal.

For melting NaF crystal, large amount of energy is required. Hence, its m.p is higher than Ne, Ar and H<sub>2</sub>O.

- Diamond has network of carbon atoms, which are joined together by strong covalent bonds in a tetrahedral manner.

Due to strong network of C atoms, it is very difficult to separate these atoms from each other. Hence Diamond has higher m.p than Ne, Ar, H<sub>2</sub>O and NaF.



## Solved Exercise Chapter # 4 (Solids)

### ALP SMART SYLLABUS 2020

(C.W)

Multiple choice questions.

- Q1. (ii) Amorphous solids  
(a) have sharp melting points.  
(b) undergo clean cleavage when cut with knife.  
(c) have perfect arrangement of atoms.  
(d) can possess small regions of orderly arrangement of atoms.
- (iii) The molecules of CO<sub>2</sub> in dry ice form the  
(a) ionic crystals  
(b) covalent crystals  
(c) molecular crystals  
(d) any type of crystal
- (iv) Which of the following is a pseudo solid?  
(a) CaF<sub>2</sub>  
(b) Glass  
(c) NaCl  
(d) All (C.W)
- Q2. (i) Fill in the blanks  
In a crystal lattice, the number of nearest neighbours to each atom is called the coordination number.
- (ii) There are 14 Bravais lattices.
- (iii) A pseudo solid is regarded as super cooled liquid.
- (iv) Glass may begin to crystallize by a process called annealing.
- (v) Crystalline solids which exhibit the same physical properties in all directions are called isotropes.
- (vi) The branch of science which deals with the geometry and structure of crystals is called crystallography.
- Q.4 What are solids. Give their general properties. How would you differentiate between crystalline and amorphous solids.
- Ans: Consult Page 95 on text book
- Q.5 Explain following terms with three examples.  
(i) Cleavage  
(ii) Habit of crystal  
(iii) Anisotropy  
(iv) Transition temperature  
(v) Polymorphism  
(vi) Symmetry
- Ans: Consult Page 96-97 on text book
- Q.6 Define unit cell. What are unit cell dimensions? How idea of crystal lattice developed from concept of unit cell.
- Ans: Consult Page 99 on text book

Explain the following with reasons:

- (i) Crystals showing isomorphism mostly have the same atomic ratios.
- (ii) Cubic form does not depend upon the chemical nature of atoms. It only depends on the number of atoms, sizes of atoms and way of combination of atoms.
- (iii) Crystals with same no. of atoms i.e. same atomic ratio, generally have same structure and hence they are isomorphous.
- (iv) Temperature is shown by elements having allotropic forms and showing polymorphism.
- (v) Allotropic forms of elements and polymorphous forms of a compound depend on each other by controlling temperature.
- (vi) Angles of hexagonal crystal is 120°.
- (vii) The angles of the metals decrease by increasing temperature.
- (viii) Free electrons in which positive nuclei are present at regular intervals, electrical conductivity of metal is due to the motion of free electrons.
- (ix) The electrical conductivity of metals decreases with increase in temperature.
- (x) Temperature increases in the vibrational motion of atoms.
- (xi) When water freezes, it contracts.
- (xii) Cubic structure of ice.

## CHAPTER-5 (ATOMIC STRUCTURE)

### OBJECTIVES (MCQ'S)

#### (ACCORDING TO ALP SMART SYLLABUS-2020)

- Neutron was discovered by:  
(a) Chadwick (b) C.D Anderson (c) Rutherford (d) Goldstein
- Mass of electron is:  
(a)  $9.1095 \times 10^{-31}$  kg (b)  $9.1095 \times 10^{-31}$  kg (c)  $9.1095 \times 10^{-27}$  kg (d)  $9.1095 \times 10^{-31}$  g
- Positive rays were discovered by:  
(a) J.J. Thomson (b) Rutherford (c) William Crooks (d) E. Goldstein
- $^{66}_{29}\text{Cu} \rightarrow ^{66}_{30}\text{Zn} + X$  where X is:  
(a) Proton (b) Positron (c) Electron (d) Neutron
- $^{66}_{29}\text{Cu} + ^1_0\text{n} \rightarrow ^{66}_{29}\text{Cu} + X$   
(a) Electron (b) Proton (c)  $\beta$ -rays (d)  $\gamma$ -rays
- The e/m value for the positive rays in maximum for:  
(a) Hydrogen (b) Helium (c) Oxygen (d) Nitrogen
- The positive particle produced in the discharge tube from hydrogen gas was named proton by:  
(a) Millikan (b) Goldstein (c) Rutherford (d) Chadwick
- When  $\alpha$ -particles strike on the nucleus of Be then the emitted particle is:  
(a) Proton (b) Neutron (c) Neutrino (d)  $\gamma$ radiations
- Cathode rays cause a chemical change because they have-----effect:  
(a) Oxidizing (b) Conducting (c) Reducing (d) Diffusing
- The mass of a proton is how much times more than that of an electron.  
(a) 1386 (b) 1836 (c) 6138 (d) 8136
- When fast neutron carry nuclear reaction with nitrogen it ejects particles:  
(a)  $\alpha$  (b)  $\beta$  (c)  $\gamma$  (d)  $\delta$
- The electrons in a sub shell are filled according to formula (2 times)  
(a)  $2n^2$  (b)  $2(2l+1)$  (c)  $(2l+1)$  (d) None of these
- Cathode rays can be generated at the pressure of:  
(a) 1 torr (b) 0.1 torr (c) 0.01 torr (d) 0.001 torr
- Balmer series in hydrogen spectrum lies in the region.  
(a) Ultraviolet (b) visible (c) Infra-red (d) Microwave
- Bombardment of  $\alpha$ -particles on Beryllium (Be) atoms emits neutron and this process is called:  
(a) Natural radioactivity (b) Artificial radioactivity  
(c) Pauli exclusion principle (d) Hund's rule
- Lyman series lies in:  
(a) U.V region (b) Visible region (c) I.R region (d) Microwave region
- Maximum number of electrons in an orbital is:  
(a) 6 (b) 10 (c) 14 (d) 2
- The element which has maximum number of unpaired electron is:  
(a) Cr<sub>24</sub> (b) Ca<sub>20</sub> (c) Fe<sub>26</sub> (d) Cu<sub>29</sub>
- Cathode rays strike alumina and produce a colour:  
(a) Red (b) Blue (c) Yellow (d) Green
- De Broglie equation is represented by:  
(a)  $h = \frac{mv\lambda}{\lambda}$  (b)  $m = \frac{h}{\lambda v}$  (c)  $m = \frac{h}{\lambda}$  (d)  $\lambda = \frac{h}{mv}$
- n + l value of 6d orbital is:  
(a) 08 (b) 09 (c) 10 (d) 11
- When the azimuthal quantum number is 3 then 'm' can have  
(a) 5 values (b) 7 values (c) 2 values (d) 3 values

2018

- 23- Name the electron is given by:



- (a) William Crooks (b) Stoney (c) J.J. Thomson (d) Chadwick  
 24- When one beta ( $\beta$ ) particle is emitted from the nucleus of an atom is:  
 (a) atomic number increases by 1 (b) atomic number decreases by 1  
 (c) atomic mass decreases by 1 (d) atomic mass decreases by 1  
 25- The charge on proton is:  
 (a)  $1.6022 \times 10^{-11} \text{C}$  (b)  $1.6022 \times 10^{11} \text{C}$  (c)  $1.6022 \times 10^{-19} \text{C}$  (d)  $1.6022 \times 10^{19} \text{C}$   
 26- Splitting of spectral lines when atoms are subjected to strong electric field is called:  
 (a) Compton effect (b) Zeeman effect (c) photoelectric effect (d) Stark effect

Answers														
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
A	B	D	C	D	A	C	B	A	B	C	B	B	B	B
16	17	18	19	20	21	22	23	24	25	26				
A	D	A	A	D	A	B	B	A	C	D				

## CHAPTER-5 (ATOMIC STRUCTURE) SHORT QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

1. **Discovery of Electron Proton and Neutron:**  
 1. The  $e/m$  values of positive rays obtained from hydrogen gas is 1836 times less than that of cathode rays. Justify? (7 times)  
 Ans: When we use hydrogen gas in the discharge tube, the positive rays produced which consist of only one proton. The proton is 1836 times heavier than that of electron. So its  $e/m$  value is 1836 times smaller than  $e/m$  value of electron.  
 2. How will you prove that Cathode Rays travel in Straight Lines? (2 times)  
 Ans: Cathode rays cast a shadow when an object is placed in their path. This proves that they travel in a straight line perpendicular to the surface of cathode.  
 3. Cathode rays are negatively charged? Explain it with diagram.  
 Ans: Cathode rays are negatively charged. It was proved in 1895, when J Perrin showed that when the cathode rays passed between the poles of the magnet, the path of the negatively charged particles was curved downward by magnetic field.  
 In 1897, J. Thomson established their electric charge by the application of electric field, the cathode ray particle were deflected upward. As shown in figure given below.

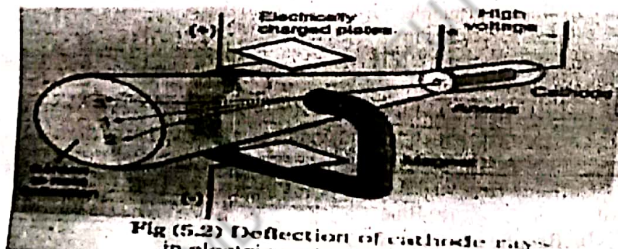
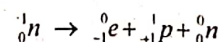


Fig (5.2) Deflection of cathode rays in electric and magnetic fields

4. Write four properties of positive rays? (8 times)  
 Ans: (i) These rays travel in a straight line in a direction opposite to the cathode rays.  
 (ii) They produce flashes on striking ZnS plate.  
 (iii) The charge to mass ratio ( $e/m$ ) for these rays is always smaller than for electrons.  
 (iv) The  $e/m$  ratio is highest when hydrogen is present.  
 5. Cathode rays can cause a chemical change. Justify?

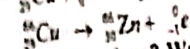
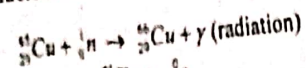
- Ans: Cathode rays can ionize the gases. The ionized gas molecules unite to form new substances. Thus, cathode rays can cause a chemical change.  
 6. Whatever gas is used in the discharge tube the nature of the cathode rays remains the same, why? (7 times) (C.W)  
 Ans: Cathode rays are actually the electrons and electrons are fundamental particles of each type of matter. Moreover, cathode rays are emitted from surface of cathode not from gas enclosed. So it is independent of nature of gas.  
 7. Why is it necessary to decrease the pressure in discharge tube to get cathode rays? (8 times) (C.W)  
 Ans: At ordinary pressure the number of particles of a gas is greater hence greater chances of collisions of cathode rays with them. These cannot reach the anode and particles becomes smaller collision decreases and conduction take place.  
 8. Justify that  $e/m$  value of positive rays for different gases are different but those for cathode rays, the  $e/m$  values are same. (2 times) (C.W)  
 Ans:  $e/m$  values of cathode rays or electrons is independent of gas of discharge tube and voltage. It means every type of substance has same type of electrons. Positive rays are produced by ionization of gas molecules. Different gases have different values of  $e/m$  of anode rays.  $e/m$  ratio of anode rays depends upon the mass of the gas. Lighter gases have high  $e/m$  values for anode rays.  
 9.  $e/m$  value of cathode ray is just equal to that of electrons. Justify? (4 times) (C.W)  
 Ans: Cathode rays are actually the electrons. Therefore their  $e/m$  value is equal to electrons.  
 10. The positive rays are also called canal rays. Explain? (7 times) (C.W)  
 Ans: Goldstein used perforated cathode in his discharge tube. These perforations are called canals. Since positive rays can pass through these canal. Hence the rays are called canal rays.  
 11. Give reason for the production of positive rays? (5 times) (C.W)  
 Ans: Positive rays are produced by the ionization of gas inside the discharge tube. Different gases have different  $e/m$  ratio. Highest  $e/m$  ratio obtained by hydrogen gas. When high velocity electrons strike with gas molecules it split up into cation and electrons. Electron move towards anode and cation move towards cathode.  
 12. Why  $e/m$  of cathode rays is equal to that of electrons? (C.W)  
 Ans: The  $e/m$  value of cathode rays shows that they are simply electrons. J.J. Thomson concluded from his experiments that cathode rays consists of stream of negatively charged particles. Stoney name these particles electrons. Thomson also determined the charge to mass ratio ( $e/m$ ) of electrons. He found that the  $e/m$  value remained the same no matter which gas was used in the discharged tube.  
 13. How the bending of cathode rays in the electric and magnetic fields shows that they are negatively charged? (C.W)  
 Ans: In 1895, J Perrin showed that when the cathode rays passed between the poles of the magnet, the path of the particles was curved downward by the magnetic field. In 1897, J. Thomson established their electric charge by the application of electric field and the particles were deflected upwards (towards the positive plate). The above mentioned two experiments showed that the cathode rays are negatively charged particles.  
 14. Describe behavior of cathode rays in magnetic field?  
 Ans: Cathode rays are negatively charged. Perrin showed that cathode rays are deflected in a magnetic field perpendicular to the lines joining the two poles.  
 15. How it was inferred that cathode rays are material particles? (2 times)  
 Ans: Cathode rays can drive a small paddle wheel placed in their path. This shows that these rays possess momentum. From this observation, it is inferred that cathode rays are not rays but material particles having a definite mass and velocity. Cathode rays are material particles and has mass  $9.1 \times 10^{-31} \text{kg}$ .  
 16. What particles are formed by the decay of free neutrons? (6 times)  
 Ans: Free neutron decay into proton, electron and neutrino. Neutrino has no charge and no mass.





Write two properties of neutrons?  
Neutrons are highly penetrating particles.  
Neutrons cannot ionize gases.

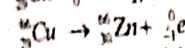
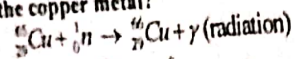
Write down nuclear reactions involved in the conversion Cu into Zn. (2 times)



What are Nuclear Reactions? Write equation for a Nuclear Reaction for the production of Neutron.

The chemical reactions involves nucleus are called nuclear reactions.

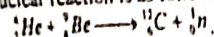
Give nuclear reactions to show the radioactive decay when a slow moving neutron hits the copper metal? (3 times)



How neutrons were discovered by Chadwick? Give the equation of nuclear reaction involved? (4 times)

Ans: Chadwick discovered neutron in 1932 and was awarded Nobel prize in Physics in 1935.

A stream of  $\alpha$ -particles produced from a polonium source was directed at beryllium ( ${}^4_2\text{Be}$ ) target. It was noticed that some penetrating radiation were produced. These radiations were called neutrons because the charge detector showed them to be neutral. The nuclear reaction is as follows.



22. How neutrons are used in treatment of cancer?

Ans: Neutrons are highly penetrating particles. Because of their intense biological effect they are being used in treatment of cancer.

23. Calculate the mass of an electron  $\frac{e}{m} = 1.7588 \times 10^{11} \text{ Coulombs Kg}^{-1}$ ? (9 times)

Ans: From e/m ratio of electrons mass of electron can be calculated as:

$$e/m = 1.7588 \times 10^{11} \text{ C/kg}$$

$$m = \frac{e}{1.7588 \times 10^{11} \text{ C/kg}} = \frac{1.6022 \times 10^{-19} \text{ C}}{1.7588 \times 10^{11} \text{ C/kg}} = 9.1096 \times 10^{-31} \text{ kg}$$

24. Differentiated between slow and fast neutron. (2 times)

Ans: When neutrons travel with an energy 1.2 MeV, they are called fast neutrons but with energy below 1eV are called slow neutrons. Slow neutrons are usually more effective than fast ones for the fission purposes.

2. Rutherford and Bohr's Model:

25. Differentiate between frequency and wave number? (4 times)

Ans: Frequency: The number of waves which passes through a given point in one second. It is denoted by  $\nu$  and units Hz.

Wave Number: It is the number of waves per unit distance. It is denoted by  $\bar{\nu}$  its units are  $\text{m}^{-1}$  or  $\text{cm}^{-1}$ .

26. Prove that  $E = h\nu$ ?

Ans: According to Planck's Quantum theory:

$$E \propto \nu \quad E = h\nu \dots (i)$$

$$\text{Where } \nu = c/\lambda \dots (ii)$$

Putting the value of  $\nu$

$$E = \frac{hc}{\lambda}$$

$$\text{Where } \bar{\nu} = \frac{1}{\lambda}$$

$$E = hc\bar{\nu}$$

27. Give two postulates of plank's quantum theory? (3 times)

Ans: (1). Energy is not emitted or absorbed continuously. Rather, it is emitted or absorbed in a discontinuous manner and in the form of wave packets. Each wave packet or quantum is associated with a definite amount of energy. In case of light the quantum of energy is often called photon.  
(2). The amount of energy associated with a quantum of radiation is proportional to the frequency ( $\nu$ ) of the radiation. Frequency is the number of waves passing through a point per second.

$$E \propto \nu \quad E = h\nu$$

Where 'h' is a constant known as Planck's constant and its value is  $6.626 \times 10^{-34} \text{ Js}$ . It is infect the ratio of energy and the frequency of a photon.

28. How do you come to know that the velocities of electrons in higher orbits are less than those of lower orbits? (4 times) (H.W)

Ans: As we known

$$r = \left( \frac{Ze^2}{4\pi\epsilon_0 m} \right) \frac{1}{\nu^2}$$

Where  $\frac{Ze^2}{4\pi\epsilon_0 m}$  is Constant

$$\text{So } r \propto \frac{1}{\nu^2}$$

Radius is inversely proportional to square of velocity. It means that if radius is increased than velocity is decreased and vice versa.

29. The radius of first orbit of hydrogen atom is  $0.529 \text{ \AA}$ . Calculate the radius of 3<sup>rd</sup> orbit of hydrogen atom?

$$\text{Ans: } r = \frac{e \cdot h^2}{\pi m e^2} \times \frac{n^2}{Z}$$

$$Z = 1$$

$$r = 0.529 \text{ \AA} \times n^2$$

$$r \propto n^2$$

$$\text{For third orbit } n = 3 \quad r = 0.529 \text{ \AA} \times (3)^2 = 4.75 \text{ \AA}$$

30. Why the potential energy of bonded electron has negative value? (H.W)

$$\text{Ans: } E_{\text{potential}} = \frac{-Ze^2}{4\pi\epsilon_0 r}$$

The minus sign indicates that the potential energy of electron decreases, when it is brought from infinity to a point at a distance 'r' from the nucleus. At infinity electron is not being attracted by anything and the potential energy of the system is zero.

31. Give two defects in Rutherford's atomic model? (20 times)

Ans: (i) A revolving electron must emit energy continuously. As a result electron will move in a spiral path and will fall into the nucleus. However it never happened.  
(ii) If electron emits energy continuously, then a continuous spectrum should be formed. Actually atoms form line spectrum.

32. Write two postulates of Bohr's atomic model? (2 times)

Ans: (i) Electrons revolve in one of the circular orbits outside the nucleus. Each orbit has fixed energy and quantum number is assigned to it.  
(ii) Electrons present in a particular orbit do not radiate energy. The energy is emitted or absorbed only, when an electron jumps from one orbit to another.

33. Give two defects in Bohr's atomic model? (5 times)

Ans: (i) The model can explain the spectrum of hydrogen and hydrogen like ions such as  $\text{He}^+$ ,  $\text{Li}^+$ ,  $\text{Be}^+$  etc. It cannot explain the spectrum of multi electron system.  
(ii) According to Bohr's model orbits are planar whereas motion of electrons takes place in three dimensional space.

(iii) This theory cannot explain Zeeman and Stark Effect.

34. Describe Stark and Zeeman affect? (8 times)



Ans: Zeeman Effect: The splitting of spectral lines of excited hydrogen atom into closely spaced lines in strong magnetic field is called Zeeman Effect.  
 Stark Effect: The splitting of spectral line of excited hydrogen atom into closely spaced lines in an electric field is called Stark Effect.

3. **Spectrum:**

35. **What is Lyman series? In which region it lies?**  
 Ans: Lyman series: When an electron jumps from higher orbit to n 1 Lyman series are obtained. It lies in ultraviolet region of spectrum.

Differentiate between line spectrum and continuous spectrum? (5 times)	
Continuous spectrum	Line spectrum
(i) In this spectrum, colors are diffused into each other and they are not separated.	It consists of dark or bright lines separated by bright or dark bands.
(ii) There is no sharp boundary between the colours.	There is a sharp boundary between the colours.
(iii) Example: Rainbow	Example: Hydrogen spectrum

37. **Differentiate between continuous and discontinuous spectrum? (2 times)**  
 Ans: Continuous spectrum: In this type of spectrum, the boundary line between the colours cannot be marked. The colours diffuse into each other. One colour merges into another without any dark space. The best example of continuous spectrum is rainbow. It is obtained from the light emitted by the Sun or incandescent (electric light) solids. It is the characteristic of matter in bulk.  
 Discontinuous spectrum: When an element or its compound is volatilized on a flame and the light emitted is seen through a spectrometer, we see distinct lines separated by dark spaces. This type of spectrum is called line spectrum or discontinuous spectrum.

38. **Differentiate between atomic emission and atomic absorption spectrum? (4 times)**

Atomic emission spectrum	Atomic absorption spectrum
In this emission bright lines are separated by dark bands. It is formed when the substance is in excited state. During its formation electrons jumps from higher to lower level.	In this emission dark lines are separated by bright band. It is formed when the substance is in unexcited state. During its formation electrons jumps from lower to higher level.

39. **How a polychromatic light split into various bands of light while passing the transparent glass prism?**

Ans: When a radiation of light is passed through a prism, the radiation undergoes refraction or bending. The extent of bending depends upon the wavelength of the photons. A radiation of longer wavelength is bent to a smaller degree than the radiation of a shorter wavelength. Ordinary, white light of radiation of all wavelengths, and after passing through the prism, white light is splitted up into radiation of different wavelengths.

40. **What is H<sub>α</sub> -line in hydrogen spectrum? Which effects explain these lines?**

Ans: H<sub>α</sub> - line in Balmer series consists of five component lines. This is called fine structure or multiple line structure of spectrum. Bohr's theory cannot explain this fine structure. Splitting of lines shows that only one quantum number is not sufficient to explain the origin of spectral lines.

41. **Write names of spectral series of hydrogen spectrum? (2 times)**

Ans: The spectrum series of hydrogen spectrum can be classified into five groups.  
 (i) Lyman series (U.V region) (ii) Balmer series (visible region)  
 (iii) Paschen series (I.R region) (iv) Brackett series (I.R region)  
 (v) Pfund series (I.R region)

42. **What is origin of Hydrogen Spectrum?**

Ans: According to Bohr, electrons in hydrogen atom may revolve in any orbit depending upon its energy. When hydrogen gas is subjected to an electric discharge, its electron move from one of the lower orbit to higher orbit, absorbing particular wavelength of energy. When it come back, the same energy is released. This energy is observed as radiation in certain region of the emission spectrum of hydrogen gas.

43. **How Davison and Germer proved dual nature of matter? (3 times)**

Ans: Two scientists Davison and Germer proved dual nature of electrons experimentally. Electrons were produced from heated tungsten filament and accelerated by applying the potential difference through charged plates. Davison and germer proved that the accelerated electrons undergo diffraction, like waves, when they fall on a nickel crystal. In this ways, wave nature of electron got verified.  
**Electron has dual nature, Justify? (2 times)**  
 Ans: According to de-Broglie wave particle concept all matter particles like electron, proton etc in motion possess the characteristics of both the particle and a wave. This is called dual nature of matter.

$$\text{Equation: } \lambda = \frac{h}{mv}$$

$\lambda$  is wave length and  $mv$  is momentum of moving matter particle.

45. **How X - rays are produced?**

Ans: X-rays are produced when rapidly moving electrons collide with heavy metals anode in the discharge tube. Energy is released in the form of electromagnetic waves when the electrons are suddenly stopped.

5. **Heisenberg's Uncertainty Principle:**

46. **Define Heisenberg's principle of uncertainty and give its mathematical expression? (10 times)**

Ans: Heisenberg's uncertainty principle: It is impossible to determine simultaneously and precisely both position and momentum of a small fast moving particle e.g. electron.

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

47. **Compare orbit and orbital? (3 times)**

	Orbit	Orbital
(i)	It is the circular path on which electrons revolve around the nucleus.	It is the region in space in which probability of finding electron is maximum.
(ii)	It is two dimensional.	It is three dimensional.

2018

48. **Why does the size of He<sup>+</sup> is much smaller than H-atom although both H-atom and He<sup>+</sup> ion are mono-electronic systems?**

Ans: It is because the atomic number (number of protons) is higher in Helium nad lesser in hydrogen. Radius or size of atom or ion depends inversely on its atomic number as follows;

$$r = \frac{\epsilon_0 n^2 h^2}{\pi m Z e^2}$$

49. **Derive de-Broglie's equation.**

Ans: According to Planck's equation

$$E = h\nu \quad \dots\dots(2)$$

According to Einstein's mass energy relationship

$$E = mc^2 \quad \dots\dots(3)$$

Where 'm' is the mass of the material particle which has to convert itself into a photon, 'and c' is the velocity of photon. Equation two values of energy;

$$h\nu = mc^2$$

Since

$$v = \frac{c}{\lambda}$$

$$\frac{hc}{\lambda} = mc^2$$

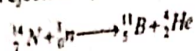
$$\lambda = \frac{h}{mv} \quad \dots\dots(1)$$



Here  $\lambda$  = de-Broglie's wavelength,  
 $m$  = mass of the particle  
 $v$  = velocity of electron

2019

50. How  ${}^4_7\text{N}$  is converted to  ${}^4_7\text{B}$ . Give equation.  
 Ans: A fast neutron ejects  $\alpha$  - particle from nucleus of N - atom and boron is produced



51. Justify that angular momentum of an electron is quantized.  
 Ans: According to Bohr's atomic model, electron can revolve only in those orbits with fixed value of angular momentum ( $mvr$ ). It is integral multiple of factor

$\frac{h}{2\pi} \cdot \frac{h}{\pi} \cdot \frac{3h}{2\pi}$  ..... The electron is bound to move in one of these orbits. So angular momentum of electron is quantized.

52. Justify that  $e/m$  value of H gas is maximum.  
 Ans: Lighter the gas used in discharge tube, larger the value of  $e/m$ . The positive particle obtained from hydrogen has smallest mass among all gases, so it has maximum  $e/m$  value. This particle is called proton.

53. Prove that  $E = hc/\lambda$

Ans: According to Plank's theory

$$E \propto \nu$$

$$E = h\nu$$

$$\nu = c/\lambda$$

$$E = hc/\lambda$$

$$\bar{\nu} = \frac{1}{\lambda}$$

$$E = hc\bar{\nu}$$

54. Justify that the distance gaps between different orbits go on increasing from lower to higher orbits.

Ans: According to Bohr's model

$$r = 0.529 \left[ \frac{n^2}{z} \right] \text{Å}^0$$

For H - atom  $z = 1$

$$\text{So } r = 0.529 \left[ \frac{n^2}{1} \right] \text{Å}^0$$

$$r = 0.529 \text{Å}^0 (n^2)$$

$$r_1 = 0.529 \text{Å}^0$$

$$r_2 = 2.1 \text{Å}^0$$

$$r_3 = 4.8 \text{Å}^0$$

$$r_4 = 8.4 \text{Å}^0$$

The order of difference between adjacent orbits are  
 $r_2 - r_1 < r_3 - r_2 < r_4 - r_3$  .....

## CHAPTER- 5 (ATOMIC STRUCTURE ) LONG QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

- Discuss properties of cathode rays? (3 times)
- How are positive rays produced in discharge tube? Give properties of these rays. (1 time)
- What is neutron? How was it discovered? Give its two properties? (1 time)
- Describe Millikan's oil drop method for determination of charge of electron? (5 times)
- Explain Rutherford's Model of Atom? (2 times)
- What is J.J Thomson's experiment for determining  $e/m$  value of electron? Evaluate mass of electron from this experiment? (4 times)
- Derive the formula for calculating the energy of an electron in  $n$ th orbit using Bohr's model? (4 times)
- Write defects in Rutherford's model of atom. How Bohr removed them? (4 times)
- Define spectrum. Explain atomic emission and atomic absorption spectrum with diagram?
- Derive the equation for radius of  $n$ th orbit of hydrogen atom using Bohr's model? (2 times)
- Give defects of Bohr's atomic model? (5 times)
- What is Atomic spectrum? Explain its types. (6 times)
- Define quantum numbers. Explain principal and magnetic quantum numbers? (2 times)
- Discuss magnetic and spin quantum numbers? (2 times)
- Discuss principal and Azimuthal Quantum numbers? (2 times)

## Solved Exercise Chapter # 5 ALP SMART SYLLABUS 2020

- Select the most suitable answer for the given one. (C.W.)
  - The nature of the positive rays depend on
    - the nature of the electrode
    - the nature of the discharge tube
    - the nature of the residual gas
    - all of the above
  - The velocity of photon is
    - independent of its wavelength
    - depends on its wavelength
    - equal to square of its amplitude
    - depends on its source
  - The wave number of the light emitted by a certain source is  $2 \times 10^6 \text{ m}^{-1}$ . The wavelength of this light will be
    - 500 nm
    - 500 m
    - 200nm
    - $5 \times 10^7 \text{ m}$
  - Rutherford's model of atom failed because
    - the atom did not have a nucleus and electrons
    - it did not account for the attraction between protons and neutrons
    - it did not account for the stability of the atom
    - there is actually no space between the nucleus and the electrons
  - Bohr model of atom is contradicted by
    - Planck's quantum theory
    - dual nature of matter
    - Heisenberg's uncertainty principle
    - all of the above
  - Splitting of spectral lines when atoms are subjected to strong electric field is called,
    - Zeeman effect
    - Stark effect
    - Photoelectric effect
    - Compton effect
  - In the ground state of an atom, the electron is present
    - in the nucleus
    - in the second shell
    - nearest to the nucleus
    - farthest from the nucleus



- (viii) Quantum number values for 2p orbitals are  
 (a)  $n = 2, l = 1$  (b)  $n = 1, l = 2$  (c)  $n = 1, l = 0$  (d)  $n = 2, l = 0$
- (ix) Orbitals having same energy are called  
 (a) hybrid orbitals (b) valence orbitals (c) degenerate orbitals (d) d-orbitals  
 (C.W)
- Q2. Fill in the blanks with suitable words.  
 (i)  $\beta$  -particles are nothing but electrons moving with a very high speed.  
 (ii) The charge on one mole of electrons is 96500 coulombs.  
 (iii) The mass of hydrogen atom is  $1.66 \times 10^{-24}$  grams.  
 (iv) The mass of one mole of electrons is  $5.48 \times 10^{-4}$ .  
 (v) Energy is released when electron jumps from higher to a lower orbit.  
 (vi) The ionization energy of hydrogen atom can be calculated from Bohr's model.  
 (vii) For d-subshell, the azimuthal quantum number has value of two.  
 (viii) The number of electrons in a given subshell is given by formula  $2(2l + 1)$ .  
 (C.W)
- Q3. Indicate true or false as the case may be.  
 (i) A neutron is slightly lighter particle than a proton. (True)  
 (ii) A photon is the massless bundle of energy but has momentum. (True)  
 (iii) The unit of Rydberg constant is the reciprocal of unit of length. (False)  
 (iv) The actual isotopic mass is a whole number. (False)  
 (v) The actual isotopic mass is applicable to macroscopic bodies. (True)  
 (vi) Heisenberg's uncertainty principle is applicable to macroscopic bodies. (True)  
 (vii) The nodal plane in an orbital is the plane of zero electron density. (True)  
 (viii) The number of orbitals present in a sublevel is given by the formula  $(2l + 1)$ . (True)  
 (ix) The magnetic quantum number was introduced to explain Zeeman and Stark effect. (True)  
 (x) Spin quantum number tells us the direction of spin of electron around the nucleus. (False)
- Q8. Derive the formula for calculating the energy of an electron in nth orbit using Bohr's model. Keeping in view this formula explain the following: (H.W)
- Ans: Consult Page 128-129 on text book  
 (a) The potential energy of the bounded electron is negative.  
 (b) Total energy of the bounded electron is also negative.
- Ans:  $E_n = -2.178 \times 10^{-18} \left[ \frac{Z^2}{n^2} \right] J$
- The negative sign shows that electron is bound by the nucleus i.e. electron is under force of attraction of nucleus.
- (c) Energy of an electron is inversely proportional to  $n^2$ , but energy of higher orbit are always greater than those of the lower orbits.
- Ans:  $E_n = -2.178 \times 10^{-18} \left[ \frac{Z^2}{n^2} \right] J$
- $E_n \propto \frac{1}{n^2}$  Where n is orbit number  
 Hence energy of electron is inversely proportional to negative value of  $n^2$  so it increases with increase in orbit no.  
 The negative sign shows that electron is bound by the nucleus i.e. electron is under force of attraction of nucleus.
- (d) The energy difference between adjacent levels goes on decreasing sharply.
- Ans: Consult Text Book Page 130
- Q11. (a) Hydrogen atom and He<sup>+</sup> are mono-electronic system, but the size of He<sup>+</sup> is much smaller than H<sup>+</sup>, why? (H.W)

- Ans: Both hydrogen atom and He<sup>+</sup> ion have one electron in their outermost shell. However, the nucleus of He<sup>+</sup> has greater positive charge (due to two protons) than that of hydrogen atom (due to one proton). Therefore, nucleus of He<sup>+</sup> attracts its electron more powerfully, as compare to hydrogen. Hence, size of He<sup>+</sup> becomes smaller than hydrogen.
- (b) Do you think that the size of Li<sup>+</sup> is even smaller than He<sup>+</sup>? Justify with calculations.  
 Both He<sup>+</sup> and Li<sup>2+</sup> ion have one electron in their outermost shell. However, the nucleus of Li<sup>2+</sup> has greater positive charge (due to three protons) than that of He<sup>+</sup> ion (due to two protons). Therefore, nucleus of Li<sup>2+</sup> attracts its electron more powerfully as compared to He<sup>+</sup>. Hence size of Li<sup>2+</sup> becomes smaller than hydrogen. A photon of light with energy  $10^{-19} J$  is emitted by a source of light.
- Q17. Convert this energy into the wavelength, frequency and wave number of the photon in terms of meters, hertz and  $m^{-1}$ , respectively.
- Ans:  $E = 10^{-19} J$ ,  $h = 6.625 \times 10^{-34} Js$   
 $C = 3 \times 10^8 m/s$ ,  $v = ?$   
 $\lambda = ?$ ,  $v = \lambda$   
 Since  $E = hv$
- Or  $v = \frac{E}{h} = \frac{10^{-19}}{6.625 \times 10^{-34}} = 1.509 \times 10^{14} s^{-1}$   
 $\lambda = \frac{c}{v} = \frac{3 \times 10^8}{1.509 \times 10^{14}} = 1.988 \times 10^{-6} m$   
 $\bar{\nu} = \frac{1}{\lambda} = \frac{1}{1.988 \times 10^{-6}} = 5.030 \times 10^5 m^{-1}$
- (b) Convert this energy of the photon into ergs and calculate the wavelength in cm, frequency in Hz and wave number in  $cm^{-1}$ .  
 $[h = 6.626 \times 10^{-34} Js \text{ or } 6.625 \times 10^{-27} ergs, c = 3 \times 10^8 ms^{-1} \text{ or } 3 \times 10^{10} cms^{-1}]$
- Ans:  $E = 10^{-19} J = 10^{-19} \times 10^7 = 10^{-12} erg$   
 $h = 6.625 \times 10^{-34} \times 10^7 = 6.625 \times 10^{-27} ergs$   
 $C = 3 \times 10^8 m/s = 3 \times 10^{10} cm/s$   
 $\lambda = ?$ ,  $v = ?$ ,  $\bar{\nu} = ?$   
 Since  $E = hv$
- Or  $v = \frac{E}{h} = \frac{10^{-12}}{6.625 \times 10^{-27}} = 1.509 \times 10^{14} s^{-1}$   
 $\lambda = \frac{c}{v} = \frac{3 \times 10^{10}}{1.509 \times 10^{14}} = 1.988 \times 10^{-4} m$   
 $\bar{\nu} = \frac{1}{\lambda} = \frac{1}{1.988 \times 10^{-4}} = 5.030 \times 10^3 m^{-1}$
- Q 19. Bohr's equation for the radius of nth orbit of electron in hydrogen atom is (C.W)
- $$r_n = \frac{\epsilon_0 n^2 h^2}{\pi m e}$$
- (a) When the electron moves from  $n = 1$  to  $n = 2$ , how much does the radius of the orbit increases.
- Ans:  $\epsilon_0 = 8.85 \times 10^{-12} C^2 J^{-1} m^{-1}$   
 $h = 6.625 \times 10^{-34} Js$   
 $m = 9.1 \times 10^{-31} Kg$   
 $e = 1.6022 \times 10^{-19} C$   
 Radius of nth orbit is given by



$$r_n = \frac{e_0 h^2}{\pi m e} \times n^2$$

$$r_n = \frac{(8.85 \times 10^{-12} \text{ C}^2 \text{ J}^{-1} \text{ m}^{-1}) (6.625 \times 10^{-34} \text{ Js})^2}{3.14 \times (9.1 \times 10^{-31} \text{ Kg}) (1.6022 \times 10^{-19} \text{ C})^2} \times n^2$$

$$r_n = 0.529 \times 10^{-10} \text{ m} \times n^2 = 0.529 \text{ \AA} \times n^2$$

Thus for  $n = 1$

$$r_1 = 0.529 \times 1^2 = 0.529 \text{ \AA}$$

Thus for  $n = 2$

$$r_2 = 0.529 \times 2^2 = 2.11 \text{ \AA}$$

$$r_3 = 0.529 \times 3^2 = 4.75 \text{ \AA}$$

Hence increase in radius =  $r_2 - r_1 = 2.11 \text{ \AA} - 0.529 \text{ \AA} = 1.581 \text{ \AA}$

(b) What is the distance travelled by the electron when it goes from  $n=2$  to  $n=3$  and  $n=9$  to  $n=10$ ?

$$\text{Sol: } r_n = 0.529 \times 10^{-10} \text{ m} \times n^2 = 0.529 \text{ \AA} \times n^2$$

Thus for  $n = 2$

$$r_2 = 0.529 \times 2^2 = 2.11 \text{ \AA}$$

Thus for  $n = 3$

$$r_3 = 0.529 \times 3^2 = 4.75 \text{ \AA}$$

$$\text{Hence distance travelled} = r_3 - r_2 = 4.75 \text{ \AA} - 2.11 \text{ \AA} = 2.64 \text{ \AA}$$

Thus for  $n = 9$

$$r_9 = 0.529 \times 9^2 = 42.849 \text{ \AA}$$

Thus for  $n = 10$

$$r_{10} = 0.529 \times 10^2 = 52.9 \text{ \AA}$$

$$\text{Hence distance travelled} = r_{10} - r_9 = 52.9 \text{ \AA} - 42.849 \text{ \AA} = 10.05 \text{ \AA}$$

Q 23 Calculate the wave number of the photon when the electron jumps from (C.W)

(i)  $n = 5$  to  $n = 2$ .

(ii)  $n = 5$  to  $n = 1$

In which series of spectral lines and spectral regions these photons will appear.

$$\text{Sol: Rydberg constant} = R = 1.097 \times 10^7 \text{ m}^{-1}$$

When electron jumps from  $n = 5$  to  $n = 2$

The wave number of the photon is given by the eq.

$$\bar{\nu} = R \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1.097 \times 10^7 \left( \frac{1}{2^2} - \frac{1}{5^2} \right)$$

$$\bar{\nu} = 1.097 \times 10^7 \left( \frac{1}{4} - \frac{1}{25} \right) = 1.097 \times 10^7 \times \frac{21}{100} = 2.30 \times 10^6 \text{ m}^{-1}$$

This spectral line is present in visible region (Balmer Series)

When electron jumps from  $n = 5$  to  $n = 1$

The wave number of the photon is given by the eq.

$$\bar{\nu} = R \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1.097 \times 10^7 \left( \frac{1}{1^2} - \frac{1}{5^2} \right)$$

$$\bar{\nu} = 1.097 \times 10^7 \left( \frac{1}{1} - \frac{1}{25} \right) = 1.097 \times 10^7 \times \frac{24}{25} = 1.05 \times 10^7 \text{ m}^{-1}$$

This spectral line is present in UV region (Lyman Series)

Q 24 A photon of a wave number  $102.70 \times 10^5 \text{ m}^{-1}$  is emitted when electron jumps from higher to  $n = 1$ .

(a) Determine the number of that orbit from where the electron falls. (C.W)

Sol: Rydberg constant =  $R = 1.097 \times 10^7 \text{ m}^{-1}$

$$n_1 = 1, n_2 = ?$$

The wave number of the photon is given by the eq.

$$\bar{\nu} = R \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$$102.7 \times 10^5 = 1.097 \times 10^7 \left( \frac{1}{1^2} - \frac{1}{n_2^2} \right)$$

$$\frac{102.7 \times 10^5}{1.097 \times 10^7} = 1 - \frac{1}{n_2^2}$$

$$0.93637 = 1 - \frac{1}{n_2^2}$$

$$\frac{1}{n_2^2} = 1 - 0.93637 = 0.0636$$

$$n_2^2 = \frac{1}{0.0636}$$

$$n_2 = \sqrt{\frac{1}{0.0636}} = 3.96 \approx 4$$

(b) Indicate the name of the series to which this photon belongs.

Sol: This spectral line is present in Lyman series.

(c) If the electron will fall from higher orbit to  $n = 2$ , then calculate the wave number of the photon emitted. Why this energy difference is so small as compared to that in part (a)?

Sol: When electron jumps from  $n = 4$  to  $n = 2$

The wave number of the photon is given by the eq.

$$\bar{\nu} = R \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = 1.097 \times 10^7 \left( \frac{1}{2^2} - \frac{1}{4^2} \right)$$

$$\bar{\nu} = 1.097 \times 10^7 \left( \frac{1}{4} - \frac{1}{16} \right) = 1.097 \times 10^7 \times \frac{3}{16} = 2.05646 \times 10^6 \text{ m}^{-1}$$

Energy difference for  $n = 4$  to  $n = 1$  can be calculated by the eq.

$$\Delta E = 2.18 \times 10^{-18} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$$\Delta E = 2.18 \times 10^{-18} \left( \frac{1}{1^2} - \frac{1}{4^2} \right)$$

$$\Delta E = 2.18 \times 10^{-18} \left( \frac{1}{1} - \frac{1}{16} \right)$$

$$\Delta E = 2.18 \times 10^{-18} \times \frac{15}{16} = 2.04 \times 10^{-18} \text{ J}$$

And the energy difference for  $n = 4$  to  $n = 2$  can be calculated by the eq:



$$\Delta E = 2.18 \times 10^{-18} \left( \frac{1}{2^2} - \frac{1}{4^2} \right)$$

$$\Delta E = 2.18 \times 10^{-18} \left( \frac{1}{4} - \frac{1}{16} \right)$$

$$\Delta E = 2.18 \times 10^{-18} \times \frac{3}{16} = 4 \times 10^{-19} = 0.4 \times 10^{-18} \text{ J}$$

The energy difference in second case is small.  
It is because electron travel more distance from  $n = 4$  to  $n = 1$  than  $n = 4$  to  $n = 2$ . And since energy is directly related to the distance of the electron. Hence energy difference in second case is smaller than the first case.

Q 25. (a) What is de-Broglie's wavelength of an electron in meters travelling at half speed of light? (C.W)

Sol: Mass of electron  $= m = 9.1 \times 10^{-31} \text{ kg}$   
Velocity of light  $= c = 3 \times 10^8 \text{ m/s}$

$$\text{Velocity of electrons} = v = \frac{c}{2} = \frac{3 \times 10^8}{2} = 1.5 \times 10^8 \text{ m/s}$$

$$h = 6.625 \times 10^{-34} \text{ Js} \quad \lambda = ?$$

Wavelength of the electron is given by

$$\lambda = \frac{h}{mv} = \frac{6.625 \times 10^{-34}}{9.1 \times 10^{-31} \times 1.5 \times 10^8} = 4.85 \times 10^{-12} \text{ m} = 4.85 \text{ pm} \quad \therefore 10^{-12} \text{ m} = 1 \text{ pm}$$

(b) Convert the mass of electron into grams and velocity of light into  $\text{cm/s}$  and then calculate the wavelength of an electron in  $\text{cm}$ .

Sol:  $m = 9.1 \times 10^{-31} \text{ kg} \times 1000 \text{ g} = 9.1 \times 10^{-28} \text{ g}$   
Velocity of light  $= c = 3 \times 10^8 \text{ m/s}$

$$\text{Velocity of electrons} = v = \frac{c}{2} = \frac{3 \times 10^8}{2} = 1.5 \times 10^8 \text{ m/s}$$

$$h = 6.625 \times 10^{-34} \text{ Js} = 6.625 \times 10^{-34} \times 10^7 \text{ erg s} = 6.625 \times 10^{-27} \text{ erg s}$$

$$\lambda = ?$$

Wavelength of the electron is given by

$$\lambda = \frac{h}{mv} = \frac{6.625 \times 10^{-27}}{9.1 \times 10^{-28} \times 1.5 \times 10^{10}} = 4.85 \times 10^{-10} \text{ cm}$$

(c) Convert the wavelength of electron from meters to

(i) nm (ii)  $\text{\AA}$  (iii) pm.

$$\lambda = 4.85 \times 10^{-12} \text{ m}$$

$$\text{Sol: } 1 \text{ m} = 10^9 \text{ nm}$$

Therefore

$$\lambda = 4.85 \times 10^{-12} \text{ m} \times 10^9 \text{ nm} = 4.85 \times 10^{-3} \text{ nm}$$

$$\text{(ii)} \quad 1 \text{ m} = 10^{10} \text{ \AA}$$

Therefore

$$\lambda = 4.85 \times 10^{-12} \text{ m} \times 10^{10} \text{ \AA} = 4.85 \times 10^{-2} \text{ \AA} = 0.0485 \text{ \AA}$$

$$\text{(iii)} \quad 1 \text{ m} = 10^{12} \text{ pm}$$

$$\lambda = 4.85 \times 10^{-12} \text{ m} \times 10^{12} \text{ pm} = 4.85 \text{ pm}$$

## CHAPTER -6 (CHEMICAL BONDING)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS-2020)

- Octet rule is not obeyed during its formation: (4 times)  
(a)  $\text{NF}_3$  (b)  $\text{CF}_4$  (c)  $\text{PCl}_5$  (d)  $\text{CCl}_4$
- Which of the following Molecule Obey Octet Rule:  
(a)  $\text{BF}_3$  (b)  $\text{BCl}_3$  (c)  $\text{NH}_3$  (d)  $\text{SF}_6$
- Noble gases are highly stable and least reactive because: (1 time)  
(a) They are gasses (b) Their valence shell are complete  
(c) They are present in zero group (d) They are very safe
- The radius of  $\text{Na}^+$  ion is:  
(a) 92 pm (b) 93 pm (c) 94 pm (d) 95 pm
- Molecule in which the distance between two carbon atoms in the largest is:  
(a)  $\text{C}_2\text{H}_6$  (b)  $\text{C}_2\text{H}_4$  (c)  $\text{C}_2\text{H}_2$  (d)  $\text{C}_6\text{H}_6$
- Total number of bonds in  $\text{C}_2\text{H}_4$  molecule are:  
(a) Six (b) Four (c) Five (d) Eight
- The number of bond in oxygen molecule is:  
(a) One  $\sigma$  and one  $\pi$  (b) One  $\sigma$  and two  $\pi$  (c) Three sigma only (d) Two sigma only
- In methanol, bond between carbon and oxygen is:  
(a) Ionic (b) Nonpolar (c) Polar (d) Co-ordinate
- $\text{CsF}$  has ionic character  
(a) 60% (b) 80% (c) 92% (d) 100%
- The most electronegative element is: (1 time)  
(a) Nitrogen (b) Fluorine (c) Oxygen (d) Hydrogen
- The shielding effect is responsible for:  
(a) The decrease in nuclear attractive influence over the valence electrons  
(b) The increase in nuclear attractive influence over the valence electrons  
(c) The decrease repulsion between nucleus and inner electrons  
(d) The increase in attraction between nucleus and inner electrons.
- The value of third ionization energy of Mg is:  
(a)  $1450 \text{ kJ mol}^{-1}$  (b)  $7730 \text{ kJ mol}^{-1}$  (c)  $7850 \text{ kJ mol}^{-1}$  (d)  $1890 \text{ kJ mol}^{-1}$
- First Ionization Energy of Mg atom is: (2 times)  
(a)  $+738 \text{ KJ Mole}^{-1}$  (b)  $+1450 \text{ KJ Mole}^{-1}$  (c)  $-349 \text{ KJ Mole}^{-1}$  (d)  $-500 \text{ KJ Mole}^{-1}$
- Ionic and co-ordinate covalent bonds are present in:  
(a)  $\text{SO}_2$  (b)  $\text{NH}_4\text{Cl}$  (c)  $\text{C}_2\text{H}_5$  (d)  $\text{H}_2\text{O}$
- Which of the following has coordinate covalent bond?  
(a)  $\text{NH}_4\text{Cl}$  (b)  $\text{NaCl}$  (c)  $\text{HCl}$  (d)  $\text{AlCl}_3$
- The molecules which cannot form co-ordinate covalent bond with  $\text{H}^+$  ion is:  
(a)  $\text{NH}_3$  (b)  $\text{H}_2\text{O}$  (c)  $\text{PH}_3$  (d)  $\text{CH}_4$
- Who developed the VSEPR theory:  
(a) Maxell (b) Boltzmann (c) Bernoulli (d) Nyholm and Gillespie
- According to VSEPR theory, the shape of  $\text{PH}_3$  molecule is:  
(a) Trigonal Pyramidal (b) Tetragonal (c) Linear (d) Trigonal planar
- An orbital which is spherical and symmetrical is  
(a) s - orbital (b) p - orbital (c) d - orbital (d) f - orbital
- Which molecule has  $\text{sp}^2$  hybridization? (2 times)  
(a)  $\text{CH}_4$  (b)  $\text{C}_2\text{H}_4$  (c)  $\text{C}_2\text{H}_2$  (d)  $\text{C}_2\text{H}_6$
- The Carbon atom in  $\text{C}_2\text{H}_4$  is:  
(a)  $\text{sp}^3$  - hybridized (b)  $\text{sp}^2$  - hybridized (c) sp - hybridized (d)  $\text{dsp}^2$  - hybridized



21. Which of the following has linear structure?  
(a) CO<sub>2</sub> (b) NH<sub>3</sub> (c) CH<sub>4</sub> (d) H<sub>2</sub>O
22. The geometry of ethane is:  
(a) Tetrahedral (b) Trigonal planar (c) Linear (d) V-shaped
23. The bond angle in NH<sub>3</sub> molecule is:  
(a) 109.5° (b) 107.5° (c) 104.5° (d) 102°
24. The molecular shape of SO<sub>2</sub> is:  
(a) Triangular planar (b) Tetrahedral (c) Pyramidal (d) Linear
25. Shape of SnCl<sub>2</sub> is:  
(a) Linear (b) bent or angular (c) trigonal (planar) (d) Tetrahedral
26. The hybridization in ammonia molecule is:  
(a) dsp<sup>2</sup> (b) sp<sup>3</sup> (c) sp<sup>3</sup> (d) sp
27. Carbon atom in CH<sub>4</sub> is hybridized:  
(a) sp<sup>3</sup> (b) sp<sup>2</sup> (c) sp (d) dsp<sup>2</sup>
28. Molecular orbital theory was proposed by  
(a) Mosley (b) Werner (c) Kossel (d) Mulikan
29. Which of the following has the highest bond order?  
(a) O<sub>2</sub><sup>+1</sup> (b) O<sub>2</sub><sup>+2</sup> (c) O<sub>2</sub><sup>+1</sup> (d) O<sub>2</sub><sup>+2</sup>
30. Bond order of O<sub>2</sub><sup>+2</sup> is:  
(a) Zero (b) one (c) Two (d) Three
31. In nitrogen molecule (N<sub>2</sub>), each nitrogen atom contributes in sharing of formation of bond:  
(A) one electron (B) two electrons (C) three electrons (D) four electrons
32. Which one has highest value of ionization energy:  
(A) Be (B) C (C) O (D) F
33. Geometry of SO<sub>2</sub> molecule is:  
(A) Linear (B) Angular (C) Tetrahedral (D) Trigonal pyramidal
34. 1 Å = \_\_\_\_\_ m  
(A) 10<sup>-10</sup> (B) 10<sup>-11</sup> (C) 10<sup>-12</sup> (D) 10<sup>-13</sup>
35. Which element has highest value of electron affinity.  
(A) Fluorine (B) Chlorine (C) Bromine (D) Iodine
36. In ethyne molecule the number and nature of bonds are:  
(A) One sigma two Pi (B) Two sigma one Pi (C) Three sigma two Pi (D) Two sigma two pi
37. Which species has unpaired electrons in its molecular orbitals  
(a) B<sub>2</sub> (b) F<sub>2</sub> (c) N<sub>2</sub><sup>2+</sup> (d) O<sub>2</sub><sup>2+</sup>
38. Nature of bonds in N<sub>2</sub> molecule is:  
(a) One sigma; two pi bonds (b) Two sigma; two pi bonds (c) Two sigma; one pi bond (d) Three pi bonds
39. The type of hybridization in BeCl<sub>2</sub> is:  
(a) sp<sup>3</sup> (b) sp<sup>2</sup> (c) sp (d) dsp<sup>2</sup>

## Answers

1	2	3	4	5	6	7	8	9	10	11	12	13	14
C	C	B	D	A	A	A	C	C	B	A	B	A	B
15	16	17	18	19	20	21	22	23	24	25	26	27	28
A	D	D	A	A	B	B	A	A	B	A	B	C	A
29	30	31	32	33	34	35	36	37	38	39	40		
D	B	B	C	D	B	A	A	C	C	A	C		

## CHAPTER-6 (CHEMICAL BONDING)

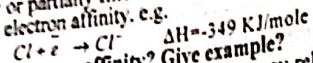
### SHORT QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

1. Causes of Chemical Combination:  
1. Define octet rule. Give two examples of compounds which do not obey this rule?  
Ans: Octet rule: The tendency of atoms to attain a maximum of eight electrons in the valence shell is known as Octet rule. Many compounds do not obey this rule. For example: PCl<sub>5</sub>, BF<sub>3</sub>, SF<sub>6</sub>, etc. (10 times)
2. What is the cause for chemical change?  
Ans: It is the natural tendency of each element in periodic table to attain the electronic configuration of nearest noble gas by:  
(i) Losing electrons (ii) Gaining electrons (iii) Sharing of electrons
3. Why the atoms of the elements other than noble gases combine with one another?  
Ans: Atoms can combine with each other to attain the 2 or 8 electrons in their valence shell. That's why they combine with each other. While noble gases have complete their octet.
4. Define bond energy with two parameters which determine its strength?  
Ans: Bond energy: The amount of energy required to break 6.02 x 10<sup>23</sup> bonds of particular type is called bond energy. Its unit is KJ/mole. It depends upon following factors.  
(i) Electronegativity difference between bonded atoms.  
(ii) Size of atoms.  
(iii) Bond length.
5. Compare bond strength of polar and non-polar molecules?  
Ans: Polar covalent bond is stronger than non-polar covalent bond. Bond energy generally increases due to increases in electronegativity difference between bonded atoms and it is because of increase in percentage ionic character. (3 times)
6. Write short note on atomic radii and ionic radii. Give example?  
Ans: Atomic radius: The one-half distance between the nuclei of two alike bonded atoms is called atomic radius. Example: Atomic radius of Na is 186 pm.  
Ionic radius: The radius of an ion while considering it spherical is called ionic radius. Example: Ionic radius of Na is 95.
7. Ionization Energy, Electron affinity and Electronegativity:  
7. Define ionization energy. Give two factors on which it depends? (4 times)  
Ans: Ionization energy: The minimum amount of energy required to remove an electron from valence shell of an isolated gaseous atom to form gaseous positive ion is called ionization energy. Factors affecting ionization energy.  
(i) Atomic radius (ii) Nuclear charge (iii) Shielding effect
8. Why atomic radius is greater than cationic radius?  
Ans: When an atom lose one or more electrons to form cation. In positive ions number of electrons are reduced but positive charge on nucleus remain same. Therefore nucleus attracts electrons inward resulting in decrease of ionic radius. Hence positive ions always have smaller size than neutral atom. (5 times)
9. Why cationic radii is smaller than anionic radii?  
Ans: The ionic radius of a cation is smaller than the atomic radius of the element from which it is derived. The ionic radius of an anion is greater than the atomic radius of the corresponding atom. This is due to the reason, that with the successive loss of electrons the nuclear charge attracts the remaining electrons with a greater force the increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valence shell electrons.
10. Why cationic radius is smaller than parent atom?  
Ans: The ionic radius of a cation is smaller than the atomic radius of the element from which it is derived. This is due to the reason that with the successive loss of electrons the nuclear charge attracts the remaining electrons with a greater force. (5 times)



11. Define electronegativity and electron affinity of an atom? (1 time)  
 Ans: **Electronegativity:** The tendency of an atom to attract shared pair of electron toward itself is called electronegativity.  
**Electron Affinity:** The amount of energy released when an electron is added to an empty or partially filled orbital of an atom in its valence shell to form an anion is called electron affinity. e.g.



12. Define electron affinity? Give example?  
 Ans: **Electron Affinity:** The amount of energy released when an electron added to an empty or partially filled orbital of an atom in its valence shell to form an anion is called electron affinity. e.g.  $\text{Cl} + e \rightarrow \text{Cl}^- \quad \Delta H = -349 \text{ KJ/mol}$  (2 times)
13. Why anionic radius is larger than atomic radius?  
 Ans: Negative ion is always larger in size than the neutral atom from which it is formed. It is because electron electron repulsions increase due to increases in number of electrons in valence shell and hold of nucleus to greater valence electrons decreases. For example:

	Atomic radius
Cl	99 pm
Cl <sup>-</sup>	181 pm

14. Why the atomic radii of the atoms cannot be determined precisely? (4 times)  
 Ans: Atomic radii cannot be determined precisely because

(i) There is no sharp boundary of an atom. The probability of finding an electron never zero even at large distance.

(ii) The electronic probability distribution is affected by neighboring atoms. (2 times)

15. Why the atomic radii increases down the group?  
 Ans: The increase in atomic radii in a group is due the increase in the number of shells and the screening effect.

16. Why the atomic radius decreases along period and increases along group?  
 Ans: The atomic radii decreases from left to the right in a period and increases from top to bottom in a group of the periodic table. The decreasing trend in a period is due to the increase in the nuclear charge. As the nuclear charge increases, the pull on the electrons is increased and size of an atom decreases. Moreover, the shielding effect remains the same from left to right in a period. The increase in atomic radii in a group is due to increase in the number of shells and the screening effect.

17. The size chlorine (Cl) atom is smaller than Cl<sup>-</sup> ion. Justify it? (2 times)  
 Ans: The ionic radius of an anion is greater than the atomic radius of the corresponding atom. The increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valence shell electrons. Hence the size of Cl<sup>-</sup> ion increases from 99 pm to 181 pm.

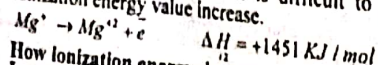
18. Define electronegativity. Give its trend in periodic table? (7 times)  
 Ans: **Electronegativity:** The tendency of an atom to attract the shared pair of electron toward itself in a molecule is called electronegativity.

Trend in periodic table:

Electronegativity values increase from left to right in a periodic table due to decrease in atomic size. Electronegativity value decrease down the group.

19. The size of anion is greater than its parent atom. Justify. (2 times)  
 Ans: The increase in the size of anion is due to the increase in electron-electron repulsion because of the increase in the valence shell electrons.
20. 1<sup>st</sup> ionization energy of Mg is lower than the 2<sup>nd</sup> ionization energy. Explain?  
 Ans:  $\text{Mg} \rightarrow \text{Mg}^+ + e \quad \Delta H = +738 \text{ KJ/mol}$

It is because due to removal of electron, hold of the nucleus on the remaining electrons increases. Hence it is difficult to remove second electron and thus ionization energy value increase.



21. How ionization energy does vary in periodic table?  
 Ans: In groups: Ionization energy decreases down the group although nuclear charge increases. It is due to increase in number of shells down the group. (9 times)

In periods: Ionization energy increase along the periods from left to right. It is due to number of shell remain same and nuclear charge increases so ionization energy also increases.

22. Ionization energy decreases down the group, although atomic number is increased. Explain it? (8 times)

Ans: It is due to increase in number of shells down the group. Due to addition of shells electrons become away from nucleus and removal of electrons become easy so ionization energy decrease down the group.

23. Why the ionic radius of Cl<sup>-</sup> ion increases from 99 pm to 181 pm? (2 times)  
 Ans: The radius of negative ion is always larger than parent atom. It is because electron-electron repulsions. Hence the size of Cl<sup>-</sup> ion increases from 99 pm to 181 pm.

24. Electronegativity increase from left to right in the periodic table. Give reason?  
 Ans: Electronegativity values increases from left to right in a periodic table due to decreased in atomic size. As atomic size decreases attraction for shared electron pair increases. Hence electronegativity value increases.

25. How electronegativity values of elements help us in understanding the nature of chemical bond?  
 Ans: Difference in electronegativity values is an index to the polar nature of bond in a molecule. If electronegativity difference is zero, the bond is non-polar. All the bonds between different atoms are polar. If difference of electronegativity value is greater than 1.7 the bond is ionic.

26. Ionization Energy is an Index to the metallic character. Justify it. (3 times)  
 Ans: Ionization Energy is an Index to the metallic character. The elements having low ionization energies are metals and those having high ionization energies are non-metals. Those with intermediate values are metalloids.

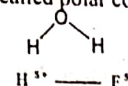
27. Define shielding effect and how it varies along period.  
 Ans: The decrease in force of attraction between nucleus and valence electron due to presence of inner shell electrons or intervening electrons is called shielding effect. Shielding effect remains almost same in a period because no new shells increase between and valence shell electron.

3. **Types of Bonds:**  
 28. Define Covalent bond. Give two example  
 Ans: According to Lewis and Kossel, a covalent bond is formed by mutual sharing of electrons between two atoms. For example, H<sub>2</sub> or Cl<sub>2</sub> molecules.

29. Define ionic and covalent radii.  
 Ans: **Ionic Radii:** The ionic radius of an ion is the radius of the ion while considering it to spherical in shape. For example ionic radius of Na<sup>+</sup> ion is 95 pm. (2 times 2018)  
**Covalent Radii:** The covalent radius of an element is defined as half of the single length between two similar atoms covalently bonded in a molecule. For example covalent radius of hydrogen is 37.7 pm.

30. Why the radius of an atom can not be determined precisely?  
 Ans: (i) There is no sharp boundary of an atom. The probability of finding an electron never becomes exactly zero even at large distance from the nucleus.  
 (ii) The electronic probability distribution is affected by neighbouring atoms. For this reason, the size of an atom may change from one compound to another.

31. Define polar covalent bond with two examples?  
 Ans: **Polar covalent bond:** The covalent bond in which bonding electrons are not shared equally by atoms is called polar covalent bond. e.g. Water and Hydrogen Fluoride.



32. How coordinate covalent bond differ from covalent bond? (3 times)  
 Ans: **Covalent bond:** The bond formed by mutual sharing of electrons between two atoms is called covalent bond. While in coordinate covalent bond shared pair of electron is donated by only one of the bonded atoms. But in some cases after bond formation no distinction is present between a coordinate covalent bond and covalent bond. e.g. in H<sub>3</sub>O<sup>+</sup>

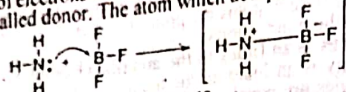
33. The distinction between a coordinate covalent bond and a covalent bond vanishes after bond formation in NH<sub>4</sub><sup>+</sup>, H<sub>3</sub>O<sup>+</sup> and CH<sub>3</sub>NH<sub>3</sub><sup>+</sup>? (2 times)  
 Ans: H<sub>3</sub>O<sup>+</sup>: Oxygen of water give its lone pair to H<sup>+</sup> to form coordinate covalent bond. The new bond is formed with same specie (H<sup>+</sup>) as two of them are already bonded. Hence all three bonds have equal status. Every bond has 33% coordinate covalent character and 66% covalent character.



$\text{NH}_3$ : The coordinate covalent bond is formed with same category ( $\text{H}^+$ ) as already discussed. Every bond has 25% coordinate covalent bond character and 75% covalent bond character.

$\text{CH}_3\text{NH}_2$ : All bonds behave alike due to same type of atoms.

34. What is coordinate covalent bond? Give example? (13 times)  
 Ans: Co-ordinate covalent bond: A bond which is formed between two atoms in which the shared pair of electrons is donated by one of the bonded atoms. The atoms which donate pair is called donor. The atom which accepts pair is called acceptor.



35. Explain the limitations of Lewis model?

Ans: Limitations of Lewis model:  
 Lewis model does not account for:

- Shapes of molecules.
- Molecular Geometries, bond polarities, bond distances.
- Various energy transitions as evident by spectroscopic techniques.

36. Lone pair of electrons on an atom occupies more space because it is under the influence of only one nucleus while bond pair are occupied of two atoms nucleus.  
 Ans: Lone pair of electrons on an atom occupies more space because it is under the influence of only one nucleus while bond pair are occupied of two atoms nucleus.

37. In  $\text{NH}_3$ , the bond angle is  $107.3^\circ$  but in  $\text{NF}_3$ , the bond angle is  $102^\circ$ . Explain?  
 Ans:  $\text{NH}_3$ : In  $\text{NH}_3$ , N is surrounded by four electron pairs. Three are bond pairs and one is lone pair. The lone pair present on N exerts greater repulsion on bond pair. Hence bond angle reduce from  $109.5^\circ$  to  $107.5^\circ$ . When these H-atoms are substituted by more electronegative atom like F in  $\text{NF}_3$  bond angle are further reduced. It is because of polarity of N-F bond. Hence bond angle is reduced to  $102^\circ$ .

38. Write order of repulsion between electron pairs according to VSEPR-theory? (1 time)

Ans: The repulsion between electron pairs decreases in the following order:  
 Lone pair - Lone pair > Lone pair - Bond pair > Bond pair - Bond pair

39. Draw and explain the geometry of  $\text{NH}_3$  molecule according to valence shell electron pair repulsion theory? (2 times)

Ans:

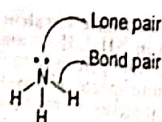


In  $\text{NH}_3$ , N is surrounded by four electron pair. Three are bond pairs and one lone pair. Due to presence of lone pair tetrahedral geometry changes to trigonal pyramidal with angle  $107.5^\circ$ .

40. Define lone pair and bond pair of electrons?

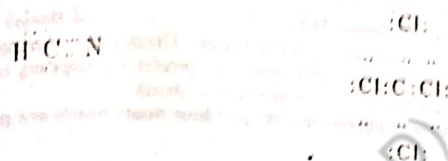
Ans: Lone pair: The electron pair which does not involve in bond formation is called lone pair.

Bond pair: The electron pair which involve in bond formation is called bond pair. For example.

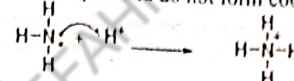


41. Write the Lewis structures for the following compounds: (2 times)  
 (a)  $\text{HCN}$  (b)  $\text{CCl}_4$

Ans:



42.  $\text{NH}_3$  can form coordinate covalent bond with  $\text{H}^+$  but  $\text{CH}_4$  not. Justify.  
 Ans:  $\text{NH}_3$  can form coordinate covalent bond with  $\text{H}^+$  due to the presence of lone pair in  $\text{NH}_3$ , while  $\text{CH}_4$  has no lone pair and do not form coordinate covalent bond.



4. Valence Shell Electron Pair Repulsion theory (VSEPR):

43. Explain the structure of  $\text{NH}_3$  molecules on the basis of VSEPR theory? (2 times)

Ans: According to the VSEPR theory in  $\text{NH}_3$  molecule, the cloud of lone pair electrons (non-bonding electrons) spread out more than that of bonding electrons. As a result, somewhat large lone pair charge cloud tend to compress the bond angles in test of the molecules.

$$N = 1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$$

The non-bonding electron in 2s orbitals take up more space and exerts a strong repulsive force on the bonding electron pairs. Consequently, to avoid a larger repulsion, the bonding electron pairs move closer that reduces the ideal bond angle form  $109.5^\circ$  to  $107.5^\circ$ . This effect compels ammonia to assume a triangular pyramidal geometry instead of tetrahedral as in methane.

44. Deviation in bond angle in water ( $\text{H}_2\text{O}$ ) molecule is more than in ammonia  $\text{NH}_3$  molecule although both have tetrahedral structures? (2 times)

Ans: There are four electron pairs in both  $\text{NH}_3$  and  $\text{H}_2\text{O}$ .

In  $\text{NH}_3$ , there are three bonding electron pairs and one lone pair. Lone pair exerts a repulsion of bonding pair due to which bond angle reduces to  $107.5^\circ$ .

While in case of  $\text{H}_2\text{O}$  there are two bonds pairs and two lone pairs on oxygen atom which exert greater repulsion and hence angle reduce from  $109.5^\circ$  to  $104.5^\circ$ .

45. Give any two postulates of valence shell electron pair repulsion theory? (1 time)

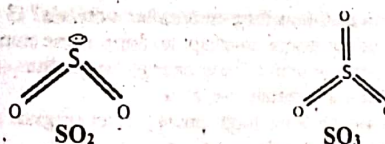
Ans: (i) Both bond pair and lone pair of electrons are involved in determining the geometry of molecules.

(ii) The electron pairs are arranged at maximum distance apart from each others to minimize repulsion.

(iii) A lone pair occupied more space than a bond pair.

46. Draw the geometry of  $\text{SO}_2$  and  $\text{SO}_3$  on the basis of VSEPR Theory.

Ans:



47. Both  $\text{NH}_3$  and  $\text{BF}_3$  are tetra atomic but different geometries why? (2 times 18)

Ans: In  $\text{NH}_3$  charge cloud of lone pair electrons (non-bonding electrons) spread out more than that of bonding electrons. As a result, some what large lone pair cloud tend to compress the bond angles in rest of the molecules. So,  $\text{NH}_3$  has trigonal bipyramidal geometry.

While in  $\text{BF}_3$  central atom B has three bonding electron pair, which are arranged at maximum distance apart at a mutual angle of  $120^\circ$ , giving a triangular planar



### 5. geometry. Valence Bond theory:

48. Write two points of valence Bond Theory? (5 times)  
 Ans: (i) The two overlapping orbitals must be valence orbitals and must be half filled.  
 (ii) Larger the overlap stronger the bond. (6 times)
49. Sigma Bond is stronger than Pi Bond. Why?  
 Ans: Strength of bond depends upon the overlapping region. Greater the overlapping, greater will be the strength of bond. Sigma bond has greater overlapping region than Pi-bond. That's why sigma bond is stronger than Pi-bond.
50. Tell the number of electron pairs shared in  $\text{CH}_4$  and how many bonds are polar?  
 Ans:



In this case four electrons paired are shared. All the bonds are polar but total molecule is non-polar due to zero dipole moment. Individual bond moment cancel each other.

51. Draw the structure of  $\text{H}_2\text{O}$  Molecule on the basis of VBT and explain it.  
 Ans:



Here, 2s and three 2p orbitals of oxygen hybridized to form four  $sp^3$  hybrid orbitals which will have a tetrahedral arrangement. Two hybrid orbitals are completely filled by the two available lone pair of electrons. The remaining two half filled hybrid orbitals undergo  $sp^3$ -s overlaps with H atoms to form two sigma bonds.

### 6. Hybridization:

52. Define orbital hybridization of orbitals? (4 times)  
 Ans: Orbital hybridization: The process in which atomic orbitals of different energy and shape are mixed together to form a new set of orbitals having same shape and energy is called orbital hybridization.

53.  $\pi$ -bonds are more diffused than sigma bonds. How? (13 times)  
 Ans: Sigma bond is formed by head on overlap of atomic orbitals of atoms. Therefore electron density is largely present between the nuclei. However Pi-bond is formed by sideways overlap of atomic orbitals. Therefore electron density is largely present above and below the plane of nuclei. Hence Pi-bond is more diffused than sigma bond.

54. Define degenerate orbitals. How degeneracy of orbitals can be lifted?

Ans: The orbitals which have same energy are called degenerate orbitals. e.g. p subshell has three degenerate orbitals. In presence of magnetic field these are oriented along x, y, z axes, and called  $p_x, p_y, p_z$ .

### 7. Molecular Orbital Theory:

55. Differentiate between bonding and anti-bonding molecular orbitals? (3 times)

Ans: B.M.O: The atomic orbitals of bonding atoms overlaps to form same number of molecular orbitals. The molecular orbitals which have energy lower than original atomic orbits is called bonding molecular orbitals. i.e.  $\sigma, \pi$

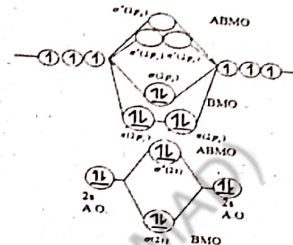
A.B.M.O: The molecular orbitals which have high energy than original atomic orbitals is called anti-bonding molecular orbitals. i.e.  $\sigma^*, \pi^*$

56. Why the energy of antibonding molecular orbital is higher than corresponding bonding molecular orbitals? (3 times)

Ans: When atomic orbitals are combine they form molecular orbitals. Some of these molecular orbitals are stable while others are unstable. Those orbitals which are stable have lower in energy than parent orbitals. And those which have higher

57. energy are unstable. Hence A.B.M.Os have higher energy than B.M.Os. Draw molecular orbital picture of  $\text{N}_2$  molecule and also calculate its bond order?

Ans:



The bond order of  $\text{N}_2$  molecule is  $\frac{6-0}{2} = \frac{6}{2} = 3$  which corresponds to the triple bond consisting of one sigma and two pi bonds.

58. What is paramagnetic character? Give the reason for paramagnetic character of oxygen? (3 times)

Ans: That property of a substance by which it is attracted towards the magnet due to the presence of unpaired electrons is called paramagnetic character. The orbitals accounts admirably for the paramagnetic properties of oxygen. This is one of the greatest successes of the molecular orbital theory. Liquid  $\text{O}_2$  is attracted towards the magnet. Anyhow, when two more electrons are given to  $\text{O}_2$ , it becomes  $\text{O}_2^{2-}$ . The paramagnetism vanished. Similarly in  $\text{O}_2^{+2}$ , the unpaired electrons are removed and paramagnetic property is no more there.

59. Discuss He-molecules on the basis of MOT?

Ans: The electronic configuration of He is  $1s^2$ . Each He atom contributes two electrons. Two electrons enter bonding molecular orbital  $\sigma(1s)$  orbitals and the remaining two electrons go to antibonding  $\sigma^*(1s)$  molecular orbital.

The bond order for  $\text{He}_2$  is zero  $\frac{2-2}{2} = \frac{0}{2} = 0$  and thus  $\text{He}_2$  molecule is not formed.

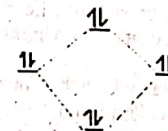
60. Represent the molecular orbitals of  $\text{N}_2$  molecule in the increasing order of energy?

Ans: The molecular orbitals of  $\text{N}_2$  in increasing order of their energy is as follows:  $\sigma(1s^2) < \sigma^*(1s^2) < \sigma(2s^2) < \sigma^*(2s^2) < \pi(2p_x^2) = \pi(2p_y^2) < \sigma(2p_z^2) < \pi^*(2p_x) < \pi^*(2p_y) < \sigma^*(2p_z)$ .  
 $B.O = \frac{6-0}{2} = \frac{6}{2} = 3$

61. Helium is diamagnetic in nature, justify?

Ans: According to MOT when electrons of Helium are distributed two electrons go into bonding molecular orbital and two into anti bonding molecular orbital. Hence all electrons are paired and He show diamagnetic behavior.

ABMO



BMO

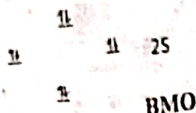
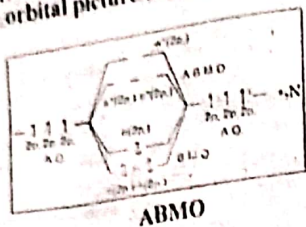
Bond order:

$$B.O = \frac{\text{No. of } \bar{e} \text{ s in B.M.O} - \text{No. of } \bar{e} \text{ s A.B.M.O}}{2}$$

$$B.O = \frac{2-2}{2} = 0$$

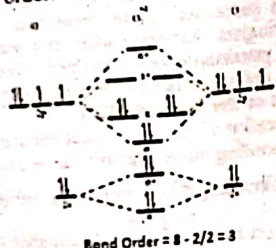


62. He has zero bond order. Sketch the molecular orbital picture of  $N_2$ ?  
Ans:



$$\text{Bond order} = \frac{6-0}{2} = \frac{6}{2} = 3$$

63. Define bond order and what is bond order of  $O_2^{2+}$ ?  
Ans: Bond order: The number of bonds formed between two atoms by overlap of atomic orbitals is called bond order.



$$\text{Bond order} = \frac{6-4}{2} = \frac{2}{2} = 1$$

64. Prove that Bond Order of  $He_2$  is zero. Or Why  $He_2$  is not possible? (2 times)  
Ans: The electronic configuration of He is  $1s^2$ . Each He atom contributes two electrons. Two electrons enter bonding molecular orbital  $\sigma(1s)$  orbitals and the remaining two electrons go to antibonding  $\sigma^*(1s)$  molecular orbital.

$$\text{The bond order for } He_2 \text{ is zero } \frac{2-2}{2} = \frac{0}{2} = 0 \text{ and thus } He_2 \text{ molecule is not formed.}$$

65. Differentiate between atomic and molecular orbital?

Ans: Atomic orbital: It is the region around the nucleus of an atom in which the probability of finding the electrons is maximum. It has no sharp boundaries and can have different shapes.

Molecular orbital: The orbital which has more than one nuclei buried in its electronic cloud is called molecular orbital.

66. Why helium gas do not exist in diatomic state? (3 times)

Ans: The electronic configuration of He is  $1s^2$ . Each He atom contributes two electrons. Two electrons enter bonding molecular orbital  $\sigma(1s)$  orbitals and the remaining two electrons go to anti bonding  $\sigma^*(1s)$  molecular orbital.

$$\text{The bond order for } He_2 \text{ is zero } \frac{2-2}{2} = \frac{0}{2} = 0 \text{ and thus } He_2 \text{ molecule is not formed.}$$

MOT is superior to VBT, explain?

(8 times)

67. Ans:

VBT	MOT
(i) According to this theory valence shell atomic orbitals of atoms overlap with each other to form covalent bond.	According to this theory atomic orbitals combine to give new orbital called molecular orbitals.
(ii) The atoms retain their individual identity.	The atoms lose their identity.
(iii) V.B.T fails to explain paramagnetic and diamagnetic behaviors.	MOT can well explain paramagnetic and diamagnetic behavior.

68. Define Covalent Radius with an example?

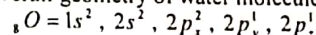
Ans: Covalent radius is defined as half of the single bond length between two similar atoms covalently bonded in a molecule. The covalent radius of hydrogen, for example is 37.7 pm. It is half of the single bond length (75.4 pm) between the two H atoms in H-H molecule.

69. Why sigma bond is stronger than Pi-bond? (2 times)

Ans: Strength of bond depends upon the overlapping region. Greater will be overlapping region, greater will be the strength of bond. Sigma bond has greater overlapping region than Pi bond, so it is stronger than Pi bond.

2018

70. Bond angle in water is  $104.5^\circ$  instead of  $109.5^\circ$ , give reason.  
Ans: Water ( $H_2O$ ), a triatomic molecule is expected to be an  $AB_2$  type linear molecule like  $BeCl_2$  and  $CO_2$ . But, experimental evidences confirm a bent or angular geometry. VSEPR theory, successfully justifies the experimental results by arguing the participation of lone pairs, in addition to bond pairs in determining overall geometry of water molecule.



Two of the corners of a tetrahedron are occupied by each of the two lone pairs and their repulsive action among themselves and on bond pairs, the bond angle in  $H_2O$  is further reduced to  $104.5^\circ$ .

71. What is the basic assumption for Valence Shell Electron Pair Repulsion (VSEPR) theory?

Ans: The valence electron pairs (lone pairs and the bond pairs) are arranged around the central atom to remain at a maximum distance apart to keep repulsions at a minimum.

72. Differentiate between polar and non-polar covalent bond. (2 times)

Ans:

Non-Polar Covalent Bonds	Polar Covalent Bonds
In such bonds, the bonding electron pairs are equally shared. For example, in $H_2$ or $Cl_2$ molecules, the two electrons forming the covalent bond are equally shared by the two identical atoms having same electronegativities. Due to an even distribution of charge, the bonded atoms remain electrically neutral. Hydrogen $H:H$ or $H-H$	When two different atoms are joined by a covalent bond, the electron pair is not equally shared between the bonded atoms. The bonding pair of electrons will be displaced towards the more electronegative atom. This would make one end of the molecule partially positive and the other partially negative as shown by the following example. Hydrogen fluoride $H^{\delta+}-F^{\delta-}$

73. 75.4 pm compromise distance between the bonded hydrogen atoms. Justify.  
Ans: The covalent radius of hydrogen, for example, is 37.7 pm. It is half of the single bond length (75.4 pm) between the two H atoms in H-H molecule, as shown in Fig.



The covalent radius of an atom can be used to determine the covalent radius of another atom. For example, the experimentally determined bond length of C-Cl in  $CH_3Cl$  is 176.7 pm. The covalent radius of Cl-atom being known as 99.4 pm, that of C-atom can be calculated by subtracting this value from C-Cl bond length. So the covalent radius of C-atom =  $176.7 - 99.4 = 77.3$  pm.



Electronegativity difference between the bonded atoms is an index to the polar nature of covalent bond. Justify.

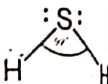
Ans: A comparison of electronegativities shows that the values increase in a period with the decrease in atomic size. These values decrease in a group as the size of the atoms increase. The electronegativity differences of the elements can be related to the properties of bonds such as dipole moments and bond energies. The difference in the electronegativity values of the bonded atoms is an index to the polar nature of the covalent bond. When the difference is zero, the bond between the two atoms is non-polar. Thus, all the bonds formed between different elements are mostly non-polar in character, while those formed from ionic bonds, polar. Elements of widely different electronegativities from ionic bonds.

75. How bond length is effected by change of hybridization state?  
Ans: Moreover, hybridization scheme involved, also explains the shortening of bonds due to the predominant participation of s-orbital. Since, the 2s-orbital of carbon has smaller mean radius than the 2p-orbitals. It would be expected that greater the s character in the hybrid orbitals used, the shorter will be the bond distance. Thus, the C-C bond lengths are 154, 133 and 120 pm for ethane, ethene and ethyne, respectively where s orbital contribution increases from  $sp^3$  to  $sp$ . Further, p-bonding also reduce the internuclear bond distance.

76. How electron affinity changes in a group?  
Ans: In groups, the atomic radii increase with the increase in the proton number due to successive increase of electronic shells. This also exerts a shielding effect on the force of attraction between the nucleus and the valence electrons. Thus, the electron affinities usually decrease from top to bottom.

2019

77. Explain geometry of  $H_2S$  molecule on the basis of VSEPR Theory.  
Ans: In valence shell of Sulphur, there are one lone and two bond pairs. They repel each other at maximum distance. As lone pair has greater repulsion with bond pairs, so geometry of  $H_2S$  is bent with bond angle  $92^\circ$ .



78. Ionic compounds are mostly soluble in water but insoluble in non-polar solvents. Give reason.  
Ans: Ionic compounds are water soluble due to ion-dipole forces. The anions are attracted by partial positive hydrogen of water and cations by partial negative oxygens.

79. The difference in E.N of bonded atoms is index of polar nature of covalent bond.  
Ans: No such interactions are developed by non-polar solvents. It means when E.N difference is zero, the bond between two atoms is non-polar, while bond formed between dissimilar atoms is mostly polar.

80. No bond in chemistry is 100% ionic.  
Ans: A difference of 1.7 units shows equal contributions of ionic and covalent bonds. Highest ionic character found in  $CsF$  is 92%. No metal is more electropositive than Cs and no non-metal is more electronegative than Fluorine, so there is no chance of 100% ionic character.

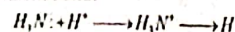
81. Define Electronegativity. Name element with highest value of E.N.  
Ans: Relative ability of an atom to attract shared electron pair towards itself is called electronegativity. Fluorine has highest E.N value according to Pauling scale. Its E.N value is 4.0.

82. Describe  $sp^2$  hybridization. Mention the molecule with  $sp^2$  hybridization.  
Ans:  $sp^2$  - Hybridization: Mixing of one s and two p orbitals to give three  $sp^2$  hybrid orbitals is called  $sp^2$  hybridization.

In  $BF_3$ , B is  $sp^2$  hybridized.  
In ethene, C is  $sp^2$  hybridized.

83.  $NH_3$  can form coordinate covalent bond with  $H^+$  ion. Explain.

Ans: There is a lone pair in valence shell of N in  $NH_3$ . It can donate to  $H^+$  ion to make coordinate covalent bond.



Ammonium ion

84. Why polar bond is stronger than non-polar bond.

Ans: Polar bond has +ve and -ve ends which provide extra force of attraction. Thus polar bond is more stronger than non-polar bond.  
85. Define bond order. Calculate bond order for  $N_2$  molecule.

Ans: Bond order =  $\frac{\text{No. of electrons in BMO's} - \text{No. of electrons in ABMO's}}{2}$   
For Nitrogen

$$\text{Bond order} = \frac{6 - 0}{2} = 3$$

86. Differentiate between Sigma & Pi-bond.

Ans: Sigma bond: When partially filled atomic orbitals overlap in linear way and probability of finding electron is on the line joining two nuclei, sigma bond is formed.

Pi-bond: When partially filled atomic orbitals overlap in parallel fashion, the probability of finding electron lies above and below the line joining two nuclei,  $\pi$ -bond is formed.

### CHAPTER-6 (CHEMICAL BONDING) LONG QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

1. Define the term electronegativity. Discuss its variation in the periodic table?  
2. What is ionization energy and discuss variation of ionization energy in periodic table. Also explain how ionization energy is an index of metallic character? (3 times)

3. Define electron affinity. Name the factors affecting it. How does it vary in the periodic table? (3 times)

4. Define covalent bond. Discuss its type with suitable examples?

5. Define coordinate covalent bond. Explain with the help of two examples?

6. Differentiate clearly between covalent and coordinate covalent bond with the help of two examples in each case?

7. Give important postulates of VSEPR theory? (5 times)

8. Write the main postulates of VSEPR theory, illustrate with the explanation of the structure of one molecule?

9. Draw the shapes of following molecules according to VSEPR theory.

i)  $BeCl_2$  ii)  $BF_3$  iii)  $NH_3$  iv)  $H_2O$

10. Write the main postulates of VSEPR theory and explain the structure of ammonia on the basis of this theory? (2 times)

11. Write down the main postulates of VSEPR theory and discuss the structure of  $NH_3$  with reference of this theory? (5 times)

12. Explain valence bond theory (VBT) with examples showing the overlapping of different orbitals?

13. How will you describe paramagnetic character of  $O_2$  molecules on basis of MOT?

14. Explain molecular orbital theory. Give molecular orbital configuration & structure of  $O_2$ ?

15. What is atomic orbital hybridization? Explain bonding and structure of ethane ( $C_2H_6$ ) on the basis of hybridization?

16. What is orbital hybridization? Explain the geometry of ethylene molecule on the basis of orbital hybridization? (4 times)

17. Explain  $sp^3$  hybridization with the help of two examples? (4 times)

18. Describe the structure of methane and ethyne on the basis of hybridization theory? (3 times)

19. Define  $sp^2$  hybridization and on its basis explain the structure of ethane? (2 times)

20. Write a note on  $sp$  hybridization and give example of ethyne. (2 times)

21. Define hybridization of atomic orbitals. Discuss  $sp^3$  hybridization in view of the structure of methane? (2 times)

22. Define orbital hybridization. Classify it and explain structure of ethane in view of orbital hybridization? (2 times)

23. Calculate the bond order of  $O_2$  molecule by making energy level diagram. Also show that it is paramagnetic. (4 times)

24. Explain important points of molecular orbital theory and draw structure of nitrogen ( $N_2$ ) molecule according to this theory? (2 times)



# Solved Exercise Chapter # 6

## ALP SMART SYLLABUS 2020

(C.W)

- Q.1 Select the correct statement  
An ionic compound A.B. is most likely to be formed when  
(i) the ionization energy of A is high and electron affinity of B is low.  
(a) the ionization energy of A is low and electron affinity of B is high.  
(b) both the ionization energy of A and electron affinity of B are high.  
(c) both the ionization energy of A and electron affinity of B are low.  
(d) both the ionization energy of A and electron affinity of B are low.
- (ii) The number of bonds in nitrogen molecule is (b) one  $\sigma$  and two  $p$   
(a) one  $\sigma$  and one  $p$   
(c) three sigma only  
(d) two  $\sigma$  and one  $p$
- (iii) Which of the following statement is not correct regarding bonding molecular orbitals?  
(a) Bonding molecular orbitals possess less energy than atomic orbitals from which they are formed.  
(b) Bonding molecular orbitals have low electron density between the two nuclei.  
(c) Every electron in the bonding molecular orbitals contributes to the attraction between atoms.  
(d) Bonding molecular orbitals are formed when the electron waves undergo constructive interference.
- (v) Which of the hydrogen halides has the highest percentage of ionic character?  
(a) HCl (b) HBr (c) HF (d) HI
- (vi) Which of the following species has unpaired electrons in antibonding molecular orbitals.  
(a)  $O_2^{+2}$  (b)  $N_2^{+2}$  (c)  $B_2$  (d)  $F_2$  (C.W)
- Q.2 Fill in the blanks  
(i) The tendency of atoms to attain maximum eight electrons in the valence shell is called completion of octet.  
(ii) The geometrical shape of  $SiCl_4$  and  $PCl_5$  can be explained on the basis of sp<sup>3</sup> hybridization.  
(iii) The VSEPR theory stands for valence shell electron pair repulsion theory.  
(iv) For  $N_2$  molecule, the energy of  $\sigma$  2p orbital is greater than  $\pi$  2p orbital.  
(v) The paramagnetic property of  $O_2$  is well explained on the basis of MO theory in terms of the presence of unpaired electrons in two MO orbitals.
- Q.3 Classify the statements as true or false. Explain with reasons. (C.W)  
(i) The core of an atom is the atom minus its valence shell. (True)  
(ii) The molecules of nitrogen ( $N \equiv N$ ) and acetylene ( $HC \equiv CH$ ) are not isoelectronic. (False)  
(iii) There are four coordinate covalent bonds in  $NH_4^+$  ion. (False)  
(iv) A  $\sigma$ -bond is stronger than a  $\pi$ -bond and the electrons of  $\sigma$ -bond are more diffused than  $\pi$ -bond. (False)  
(v) The bond energy of hetero nuclear diatomic molecules increases with the decrease in the electronegativities of the bonded atoms. (False)  
(vi) The first ionization energies of the elements rise steadily with the increasing atomic number from top to bottom in a group. (False)  
(vii) A double bond is stronger than a single bond and a triple bond is weaker than a double bond. (False)  
(ix) The bonds formed between the elements having electronegativity difference more than 1.7 are said to be covalent in nature. (False)  
(x) The repulsive force between the two bonding pairs is less than that between two lone pairs. (False)  
(xi) The number of covalent bonds an atom can form is related to the number of unpaired electrons it has. (False)  
(xii) The rules which govern the filling of electrons into the atomic orbitals also govern the filling of electrons into the molecular orbitals. (False)

What is a chemical bond? Discuss the formation of ionic and covalent bonds. How does the electronegativity differences differentiate between ionic and covalent bond? (H.W)

Q.4 Consult Text book Page # 155,162-164

Ans: (a) Define ionization energy and electron affinity. How these quantities change in the periodic table. What factors are responsible for their variation? (H.W)

Q.5 Consult Text book Page # 159-161

Ans: (b) Explain, what do you understand by the term electronegativity? Discuss its variations in the periodic table. How does it affect the bond strengths? (C.W)

Ans: Write the Lewis structures for the following compounds: (C.W)

Q.6 (i) HCN (ii)  $CCl_4$  (iii)  $CS_2$  (iv)  $H_2N \rightarrow AlF_3$   
(v)  $NH_4OH$  (vi)  $H_2SO_4$  (vii)  $H_3PO_4$  (viii)  $K_2Cr_2O_7$   
(ix)  $N_2O_5$  (x)  $Ag(NH_3)_2NO_3$

Consult Text book Page # 164-165

Ans: (a) Explain qualitatively the valence bond theory. How does it differ from molecular orbital theory? (H.W)

Valence Bond Theory	Molecular orbital Theory
According to this theory valence shell atomic orbitals of atoms overlap with each other to form covalent bonds.	According to this theory, atoms orbitals of atoms combine together to give new orbitals called molecular orbitals. This combination of orbitals is called Linear Combination of Atomic Orbitals. (LCAO)
Atomic orbitals keep their identity in the molecule.	Atomic orbitals lose their identity in the molecule.
It does not explain the paramagnetic behavior of $O_2$ molecule.	It explains the paramagnetic behavior of $O_2$ molecule.

(b) How the bonding in the following molecules can be explained with respect to valence bond theory?  $Cl_2$ ,  $O_2$ ,  $N_2$ ,  $HF$ ,  $H_2S$ .

Ans: Formation of  $Cl_2$  molecule.

$Cl_2$  molecule is formed by two chlorine atoms.

The electronic configuration of Cl is

$17Cl \quad 1s^2 2s^2 2p^6 3s^2 3p^2 3p_x^1 3p_y^1 3p_z^1$

Two Cl atoms have half-filled  $3p_z$  orbitals. These orbitals overlap to give  $Cl_2$  molecule with a single bond ( $\sigma$ -bond) between two Cl atoms.

Formation of HF molecule.

HF molecule is formed by combination of hydrogen and fluorine atoms. The electronic configuration of F & H are

$9F \quad 1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$

According to V.B.T half-filled  $2p_x$  orbital of F atom overlaps with half-filled  $1s$  orbital of H to form a  $\sigma$ -bond and thus HF molecule is formed.

Formation of  $O_2$  molecule

$O_2$  molecule is formed by combination of two oxygen atoms.

The electronic configuration of O is

$8O \quad 1s^2 2s^2 2p_x^2 2p_y^1 2p_z^1$

Two O atoms have half-filled  $2p_y$  &  $2p_z$  orbitals.  $2p_y$  orbitals of two O atoms overlap head on to give a  $\sigma$ -bond, while two  $2p_z$  orbitals of two atoms overlap to give a  $\pi$ -bond by parallel overlapping. Thus there is one  $\sigma$ -bond and one  $\pi$ -bond between two oxygen atoms in  $O_2$  molecule.

Formation of  $N_2$  molecule

$N_2$  molecule is formed by combination of two nitrogen atoms.

The electronic configuration of N is

$7N \quad 1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$



Two N atoms have half-filled  $2p_x$ ,  $2p_y$  and  $2p_z$  orbitals.  $2p_x$  orbitals of two N atoms overlap head on to give a  $\sigma$ -bond, while two  $2p_y$  orbitals and two  $2p_z$  orbitals of two atoms overlap to give a  $\pi$ -bond by parallel overlapping. Thus there is one  $\sigma$ -bond and two  $\pi$ -bonds between two nitrogen atoms in  $N_2$  molecule. The molecules  $NF_3$  and  $BF_3$  all have molecular formulae of the type  $XF_3$ . But they have different structural formulas. Keeping in view VSEPR theory sketch the shape of each molecule and explain the origin of different in shapes. (C.W.)

Ans: In  $NF_3$  nitrogen is  $sp^3$ -hybridized. There are four electron pairs present around nitrogen atom. It forms three covalent bonds with three fluorine atoms. There is one lone pair present on nitrogen atom. Hence according to VSEPR theory, it will have a trigonal pyramidal structure like  $NH_3$ . However bond angle in  $NF_3$  is  $102^\circ$  which is less than that of  $NH_3$  i.e.  $107.5^\circ$ . It is because of F is more electronegative than H. Therefore, polarity of N-F bond pulls the lone pair of N closer to the nucleus. This it exerts greater repulsions on bond pairs, hence bond angle is reduced to  $102^\circ$ .

BF<sub>3</sub>

In  $BF_3$ , B is  $sp^2$ -hybridized. It is surrounded by three electron pairs. It forms three covalent bonds with three fluorine atoms. There is no lone pair on boron. These electron pairs are present at a maximum angle of  $120^\circ$  from each other in order to have minimum repulsion between them and form trigonal planar geometry. Thus  $BF_3$  molecule is trigonal planar with bond angle of  $120^\circ$ .

Q.10 The species  $NH_3$ ,  $NH_4^+$ ,  $NH_4^-$  have bond angles of  $105^\circ$ ,  $107.5^\circ$  and  $109.5^\circ$  respectively. Justify these values by drawing their structures. (C.W.)

Ans:  $NH_4^+$ 

In  $NH_4^+$  nitrogen is  $sp^3$ -hybridized. It forms three covalent bonds and one coordinate covalent bond with hydrogen atoms. There is no lone pair present on nitrogen atom. Hence, according to VSEPR theory, it will have a regular tetrahedral structure with bond angle of  $109.5^\circ$ .

 $NH_3$ 

In  $NH_3$  nitrogen is  $sp^3$ -hybridized. It forms three covalent bonds with hydrogen atoms. There is also present a lone pair on nitrogen atom. Hence, according to VSEPR theory, it will have a trigonal pyramidal structure. Since lone pair exerts more repulsion on bond pair. Hence the bond angle will be reduced from  $109.5^\circ$  to  $107.5^\circ$ .

 $NH_2^-$ 

In  $NH_2^-$  nitrogen is  $sp^3$ -hybridized. It forms two covalent bonds with hydrogen atoms. There are also present two lone pair on nitrogen atom. Hence, according to VSEPR theory, it will have a bent structure. Since two lone pair exerts more repulsions on bond pairs. Hence the bond angle will be reduced from  $109.5^\circ$  to  $105^\circ$ .

Q.18 Explain the following with reasons: (C.W.)

(ii) The distinction between a coordinate covalent bond and a covalent bond vanishes after bond formation in  $NH_4^+$ ,  $H_3O^+$  and  $CH_3NH_3^+$ .

Ans: A co-ordinate covalent bond is formed when both the shared electrons are donated by one of the atoms.

After bond formation in many cases no distinction remains between co-ordinate covalent bond and covalent bond.

e.g. consider the formation of  $H_3O^+$  ion.

In  $H_3O^+$  ion, experimentally it has been found that all bonds are equivalent. Thus there is no distinction between co-ordinate covalent bond and covalent bond. Similarly formation of  $NH_4^+$  ion is

And formation of  $CH_3NH_3^+$  ion is

(vi) The dipole moments of  $CO_2$  and  $CS_2$  are zero, but that of  $SO_2$  is 1.61 D.

Ans:  $CO_2$  and  $CS_2$  have linear structures which are symmetrical. In this structure individual bond moments cancel the effect of each other. Hence, dipole moment becomes zero. However,  $SO_2$  has bent or angular structure in which individual bond moments do not cancel the effect of each other. Therefore, it has dipole moment.

## CHAPTER - 7 (THERMOCHEMISTRY)

### OBJECTIVES (MCQ'S)

#### (ACCORDING TO ALP SMART SYLLABUS-2020)

- Which of the following is not a state function: (3 times)  
(a) Pressure (b) Volume (c) Temperature (d) Heat
- Standard enthalpy change is measured at:  
(a) 298K (b)  $273^\circ\text{C}$  (c)  $273\text{K}$  (d)  $373\text{K}$
- For the reaction,  $H^+ + OH^-$  the change in enthalpy is called heat of  
(a) Reaction (b) Combustion (c) Solution (d) Neutralization
- The pressure of oxygen inside the bomb calorimeter is: (2 times)  
(A) 100 atm (B) 50 atm (C) 25 atm (D) 20 atm
- The unit of heat capacity are:  
(a)  $\text{kJK}^{-1}\text{mol}^{-1}$  (b)  $\text{kJK}^{-1}\text{g}^{-1}$  (c)  $\text{kJK}^{-1}$  (d)  $\text{kJK}^{-1}\text{mol}^{-2}$
- The Born-Haber cycle is the best application of .....law.  
(a) Boyle's (b) Dalton's (c) Hess's (d) Graham's
- The amount of heat absorbed when one mole of gaseous atoms are formed from the element is called enthalpy of:  
(a) Formation (b) Reaction (c) Combustion (d) Atomization

#### Answers

1	2	3	4	5	6	7
d	a	d	d	c	c	d

## CHAPTER-7 (THERMOCHEMISTRY)

### SHORT QUESTIONS

#### (ACCORDING TO ALP SMART SYLLABUS-2020)

- Spontaneous and Non Spontaneous Reactions:**  
1. Burning of a candle is a spontaneous process. Justify? (8 times)  
Ans: A reaction will also be called spontaneous process if it needs energy to start with. Burning of candle is also a spontaneous process which needs energy to start. Once the candle is made to lit with match spark. It continues to burn afterward. Therefore burning of candle is a spontaneous process.
- Is it true that a non-spontaneous process never happen in universe? Explain it. (2 times)  
Ans: No, it is not true. Some reactions which happen in the universe are non-spontaneous e.g. when there is lightening the atmospheric nitrogen and oxygen combine to make  $NO$ .
- $$N_2 + O_2 \rightleftharpoons 2NO$$
  
3. What are thermochemical reaction, give their types? (3 times)  
Ans: Thermochemical reaction: Those reaction in which energy is either evolve or absorbed during a chemical change is called thermochemical reactions.  
Two types of those reactions:  
(i) Exothermic reactions  $C + O_2 \rightarrow CO_2$   $\Delta H = -393.7 \text{ kJ/mol}$   
(ii) Endothermic reactions  $N_2 + O_2 \rightleftharpoons 2NO$   $\Delta H = +180.51 \text{ kJ/mol}^{-1}$
- System, Surrounding and State Functions:**  
4. Describe system and surrounding? (7 times)  
Ans: System: The part of universe which is under your observation is called system.



- Surrounding: Everything that is not a part of system is called surrounding. e.g. water in a glass is a system and all around is surrounding. (7 times)
5. **Define state function?**  
**Ans:** State function: It is macroscopic property of system which has definite value for initial and final states and it is independent of the path through which change takes place. For example:

Let  $T_1$  and  $T_2$  are the temperature at initial and final state then change in temperature  $\Delta T = T_2 - T_1$  volume, enthalpy, internal energy are state functions.

3. **Enthalpy:**  
**Ans:** Differentiate between internal energy and enthalpy? (7 times)  
**Internal energy:** The sum of all possible energies of atoms, ions or molecules in a system. It is not possible to measure the absolute value of internal energy of a system. However change in energy can be measured  $\Delta E = E_2 - E_1$

**Enthalpy:** Total heat content of a system or sum of internal energy and product of pressure and volume is called enthalpy.  $H = E + PV$  (4 times)

7. **Prove  $\Delta E = q_v$ ?**  
**Ans:** According to first law of thermodynamics  $\Delta E = q + w$   
 So  $\Delta E = q + P\Delta V$   
 Where  $w = P\Delta V$   
 When volume change is zero then  $\Delta V = 0$ ,  $\Delta E = q + P(0) = q_v$
8. **State first law of thermodynamics with its mathematical form? (5 times)**  
**Ans:** First law of Thermodynamic: It states that "Energy can neither be created nor be destroyed but it can change from one form to other. Or Energy of system and surrounding remains constant".  
 Mathematical form:

	$\Delta E = E_2 - E_1 = q + w$	Work = $F \times d$
Where	$\Delta E = q + w$ $w = P\Delta V$	In chemistry pressure volume work is determined
So	$\Delta E = q + P\Delta V$	Work = $P \Delta V$
When	$\Delta V = 0$ then $\Delta E = q + P(0)$	
At constant volume	$\Delta E = q_v$	

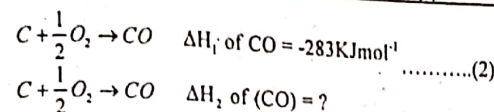
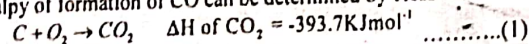
9. **What is meant by heat (q) and work (w) in thermochemistry? (3 times)**  
**Ans:** There are two fundamental ways of transferring energy to or from a system. These are heat and work. Heat is not a property of a system. It is therefore not a state function. Heat evolved or absorbed by the system is represented by a symbol q. Work is also a form in which energy is transferred from one system to another.
10. **State with one example, enthalpy of combustion? (6 times)**  
**Ans:** Enthalpy of combustion: The standard enthalpy of combustion is defined as "The enthalpy change when one mole of a substance is completely burnt in excess of oxygen under standard conditions. It is denoted by  $\Delta H_c^\circ$ ".

For example:  $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$   $\Delta H_c^\circ = -1368 \text{ kJmol}^{-1}$

11. **Define standard enthalpy of atomization with an example? (5 times)**  
**Ans:** The enthalpy change when one mole of gaseous atoms are formed from the elements under standard conditions is called standard enthalpy of atomization.

Example:  $\frac{1}{2} H_{2(g)} \rightarrow H_{(g)}$   $\Delta H_{at}^\circ = 218 \text{ kJmol}^{-1}$

12. **How enthalpy of formation of CO is determined from graphite?**  
**Ans:** Enthalpy of formation of CO can be determined by Hess's Law



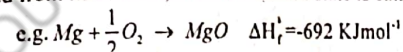
According to Hess's Law  $\Delta H = \Delta H_1 + \Delta H_2$   
 $\Delta H_2 = \Delta H - \Delta H_1 = -393.7 + 283 = -110.7 \text{ kJ/mole}$   
**Enthalpy of neutralization is for an acid and base. Explain with example? (7 times)**

13. **The amount of heat evolved when one mole of hydrogen ions ( $H^+$ ) from an acid, react completely with one mole of hydroxide ions of a base to form one mole of water is called enthalpy of neutralization.**

Example:  $NaOH + HCl \rightarrow NaCl + H_2O$   $\Delta H_n^\circ = -57.4 \text{ kJmol}^{-1}$

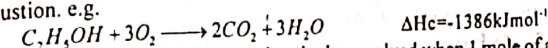
14. **Acid-base neutralization process is always exothermic. Give reason?**  
**Ans:** The standard enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions  $H^+$  from an acid, react with one mole of hydroxide ions from a base to form one mole of water. For example, the enthalpy of neutralization of sodium hydroxide by hydrochloric acid is  $-57.4 \text{ kJmol}^{-1}$ . Thus heat is evolved in acid base neutralization process is always exothermic.

15. **Define Enthalpy of formation?**  
**Ans:** Enthalpy of formation: The change in enthalpy when 1 mole of the compound is formed from its element, in their standard state is called enthalpy of formation.



16. **Enthalpy is a state function Justify?**  
**Ans:** The total heat content of a system is called enthalpy. Equation  $H = E + PV$   
 $E$ ,  $P$  and  $V$  are state functions so enthalpy is also a state function. It is not possible to measure the enthalpy of a system in a given state. However change in enthalpy can be measured for a change in the state of system.

17. **Define enthalpy of combustion and enthalpy of solution? (6 times)**  
**Ans:** Enthalpy of combustion: The amount of heat evolved when one mole of a substance is completely burnt in excess of oxygen at STP is called enthalpy of combustion. e.g.

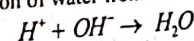


**Enthalpy of solution:** The amount of heat absorbed or evolved when 1 mole of a substance is dissolved in so much solvent that further dilution results in no detectable heat change is called enthalpy of solution.

e.g. Enthalpy of solution of  $NH_4Cl$  is  $+16.2 \text{ kJmol}^{-1}$

18. **Why the enthalpy of neutralization has the same value for any strong acid with any strong base? (2 times)**

**Ans:** When a strong acid and base are mixed together in solution form, the only change is the formation of water from  $H^+$  and  $OH^-$  ions. The net reaction is



Therefore heat of neutralization of strong base by strong acid is approximately same.

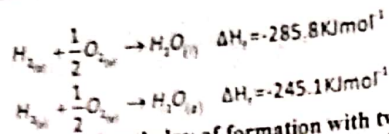
19. **Define standard enthalpy of neutralization. Give an example? (2 times)**

**Ans:** The standard enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions  $H^+$  from an acid, react with one mole of hydroxide ( $OH^-$ ) ions from a base to form one mole of water. For example, the enthalpy of neutralization of sodium hydroxide by hydrochloric acid is  $-57.4 \text{ kJmol}^{-1}$ .

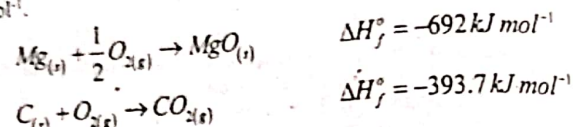
20. **Why is it necessary to mention the physical state of reactants and products in a thermochemical reactions? (6 times)**

**Ans:** Physical state of reactants and products is necessary for thermochemical reactions. Different physical states have different enthalpies. e.g the enthalpy of formation of liquid water and water vapours different from each other.

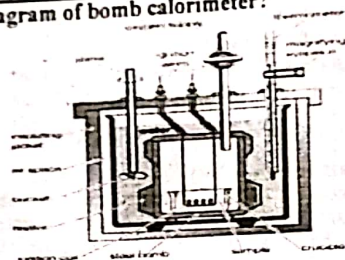




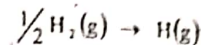
21. Define standard enthalpy of formation with two examples.  
 Ans: The standard enthalpy of formation of a compound is the amount of heat absorbed or evolved when one mole of the compound is formed from its elements. It is denoted by  $\Delta H_f^\circ$ . All the substances involved are in their standard physical states and the reaction is carried out under standard condition i.e. at 25 °C (298K) and one atm pressure. Its units are kJ mol<sup>-1</sup>.



4. **Glass & Bomb Calorimeter:**  
 22. Draw labelled diagram of bomb calorimeter? (3 times)



23. Define the term "Joule" and convert 15 calories to joules?  
 Ans: The energy expended when a force of one Newton move an object one meter in the direction in which force is applied. Where one joule is equal 10<sup>7</sup> ergs and 1 cal = 4.184 J and 15 cal = 4.184 × 15 = 62.76 J
5. **Hess's Law:**  
 24. State the Hess's law of constant heat summation? (5 times)  
 Ans: It states "If a chemical reaction takes place by different ways, the net change in energy is same regardless of the route by which the chemical change occurs. Provided the initial and final states are the same.  $\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$
25. Define Born-Haber cycle and lattice energy? (4 times)  
 Ans: Born-Haber cycle: The sum of energy changes for a closed cyclic process is zero, if the initial and final states are same.  
 Lattice energy: The amount of energy released when gaseous ions of opposite charges combine to give one mole of a crystalline ionic compound.
26. Differentiate between law of conservation of energy and Hess's law?  
 Ans: Energy can neither be created nor destroyed, but can be changed from one form to another is called Law of Conservation of Energy.  
 If a chemical change takes place by several different routes, the overall energy change is the same, regardless of the route by which the chemical change occurs, provided the initial and final conditions are the same is known as Hess's Law.
27. Explain the term "Atomization energy" with an example.  
 Ans: The standard enthalpy of atomization of an element is defined as the amount of heat absorbed when one mole of gaseous atoms are formed from the element under standard conditions. It is denoted by  $H_a^\circ$ . For example, the standard enthalpy of atomization of hydrogen is given below.

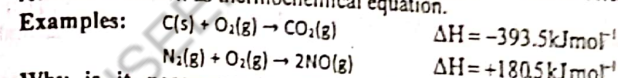


A wide range of experimental techniques, are available for determining enthalpies of atomization of elements.

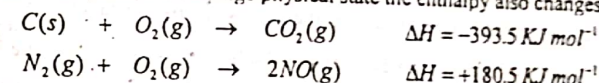
28. Differentiate between endothermic and exothermic reactions.  
 Ans:

Exothermic reactions	Endothermic reactions
A chemical reaction in which heat energy is released is called as exothermic reaction. Example: The combustion of carbon (wood or coal etc) $\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow \text{CO}_{2(g)}$ $\Delta H = -393.7 \text{ kJ mol}^{-1}$	A chemical reaction in which heat energy is absorbed is called as endothermic reaction. Example: $\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)$ $\Delta H = +180.51 \text{ kJ mol}^{-1}$

29. Define Thermochemical equation.  
 Ans: A balanced chemical equation which mentions physical states of the reactants and the products and also the amount of heat energy absorbed or released during the reaction is known as thermochemical equation.



30. Why is it necessary to mention the physical states of reactants in a thermochemical reactions?  
 Ans: A balanced chemical equation which mentions physical states of the reactants and the products and also the amount of heat energy absorbed or released during the reaction is known as thermochemical equation. Different physical states possess different enthalpies. If we change physical state the enthalpy also changes.



31. Burning of Natural gas is spontaneous process.  
 Ans: The process which proceeds on its own, is called spontaneous. But sometimes we need flame to start such process, but when it is started it proceeds on its own until completion of reaction e.g. burning of natural gas.
32. Justify that heat of formation of a compound is the sum of all other enthalpies.  
 Ans: Heat of formation of a compound  $\Delta H_f^\circ$  is net overall heat change of enthalpies. The enthalpies which make up this  $\Delta H_f^\circ$  are heat of vaporization, bond dissociation energy, ionization potential, electron affinity and lattice energy.  
 $\Delta H_f^\circ = \Delta H_s^\circ + \Delta H_d + \Delta H_{EA} + \Delta H_i + \Delta H_e$
33. What is difference between heat and temperature.  
 Ans: Heat:  
 It is a measure of total energy in a given amount of substance.  
 Its units are  $\text{KJ mol}^{-1}$  or  $\text{Kcal mol}^{-1}$   
 Temperature: It is measure of average K.E of molecules of a system.  
 It is denoted by T.  
 It is expressed in °C, °F or Kelvin scales.

## CHAPTER - 7 (THERMOCHEMISTRY)

### LONG QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

- Define enthalpy. Give its mathematical form and prove that  $\Delta H = q_p$ ? (2 times)
- Define the following with examples: ?
  - System
  - Non-spontaneous reactions
  - Surroundings
  - Endothermic reactions



3. Define the following giving an example in each:  
 (i) Enthalpy of solution (ii) Enthalpy of atomization (3 times)  
 (iii) System (iv) Surrounding (5 times)  
 (ii)  $\Delta E = q_v$ ?  
 4. Prove that (i)  $\Delta H = q_p$   
 5. State law of thermodynamics and prove that  $\Delta E = q_v$ ? (2 times)(2018)  
 6. When 2.00 moles of  $H_2$  and 1.00 mole of  $O_2$  at  $100^\circ C$  and 1100 pressure react to produce 2.00 moles of gaseous water, 482.5 kJ of energy is evolved? What are the value of (i)  $\Delta H$  (ii)  $\Delta E$  for the production of one mole of  $H_2O(g)$ ? (4 times)  
 7. Explain the term internal energy of a system and also mention how the change in internal energy of a system can occur? (2 times)(2018)  
 8. What is first law of thermodynamics? Prove that  $\Delta H = q_p$ ? (2 times)(2018)  
 9. What is the first law of thermodynamics? Give its mathematical form. (4 times)  
 10. What is Enthalpy of a reaction? How is  $\Delta H$  of a reaction measured in laboratory by glass calorimeter? (1 time)(2017 = 1 time)  
 11. What is molar heat of combustion? How it is measured by bomb calorimeter? (1 time)(2017 = 1 time)  
 12. Differentiate between: (i) exothermic and endothermic reactions (ii) spontaneous and non-spontaneous reactions. (1 time)(2017 = 1 time)  
 13. State the construction and functioning of glass calorimeter?  
 14. Neutralization of  $100\text{cm}^3$  of 0.5 M NaOH at  $25^\circ C$  with  $100\text{cm}^3$  of 0.5 M HCl at  $25^\circ C$  raised the temperature of reaction mixture to  $28^\circ C$ . Find the enthalpy of neutralization. (2 times)  
 Specific heat of water =  $4.2\text{J K}^{-1}\text{g}^{-1}$ ?  
 15. Define enthalpy of reaction. How is it measure by glass calorimeter? (6 times)  
 16. How  $\Delta H$  can be determined by using bomb calorimeter? (6 times)  
 17. Describe how the enthalpy of combustion of a substance is measured by bomb calorimeter?  
 18. Discuss the measurement of enthalpy of a reaction with the help of Bomb calorimeter? (2 times)  
 19. State and explain with an example, the Hess's law of constant heat summation? (11 times)  
 20. Define lattice energy. How Born Haber Cycle help us to calculate the lattice energy of NaCl?  
 21. One mole of methane gas is maintained at 300K and its volume is  $250\text{cm}^3$ . Calculate the pressure exerted by the gas when it is non-ideal?

### Solved Exercise Chapter # 7

### ALP SMART SYLLABUS 2020

- Q.1 Select the suitable answer from the given choices. (C.W)  
 (i) If an endothermic reaction is allowed to take place very rapidly in the air, the temperature of the surrounding air  
 (a) remains constant (b) increases (c) decreases (d) remains unchanged  
 (ii) In endothermic reactions, the heat content of the  
 (a) products is more than that of reactants  
 (b) reactants is more than that of products  
 (c) both (a) and (b) (d) reactants and products are equal  
 (iii) Calorie is equivalent to  
 (a) 0.4184 (b) 41.84 (c) 4.184 (d) 418.4  
 (iv) The change in heat energy of a chemical reaction at constant temperature and pressure is called  
 (a) enthalpy change (b) bond energy (c) heat of sublimation (d) internal energy change  
 (v) Which of the following statements is contrary to the first law of thermodynamics  
 (a) Energy can neither be created nor destroyed.  
 (b) One form of energy can be transferred into an equivalent amount of other kinds of energy.

- (c) In an adiabatic process, the work done is independent of its path.  
 (d) Continuous production of mechanical work without supplying an equivalent amount of heat is impossible.  
 (vi) For a given process, the heat changes at constant pressure ( $q_p$ ) and at constant volume ( $q_v$ ) are related to each other as  
 (a)  $q_p = q_v$  (b)  $q_p < q_v$  (c)  $q_p > q_v$  (d)  $q_p = q_v / 2$   
 (vii) For the reaction:  $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$  the change in enthalpy is called  
 (a) heat of reaction (b) heat of formation  
 (c) heat of neutralization (d) heat of combustion  
 (viii) The net heat change in a chemical reaction is same, whether it is brought about in two or more different ways in one or several steps. It is known as  
 (a) Henry's law (b) Joule's principle  
 (c) Hess's law (d) Law of conservation of energy  
 (ix) Enthalpy of neutralization of all the strong acids and strong bases has the same value because  
 (a) neutralization leads to the formation of salt and water.  
 (b) strong acids and bases are ionic substances.  
 (c) acids always give rise to  $H^+$  ions and bases always furnish  $OH^-$  ions.  
 (d) the net chemical change involve the combination of  $H^+$  and  $OH^-$  ions to form water  
 Q.2 Fill in the blanks with suitable words. (C.W)  
 (i) The substance undergoing a physical or a chemical change forms a chemical system.  
 (ii) The change in internal energy can be measured.  
 (iii) Solids which have more than one crystalline forms possess different values of heats of formation.  
 (v) A state function is a macroscopic property of a system which is independent of the path adopted to bring about that change.  
 Q.3 Indicate the true or false as the case may be. (C.W)  
 (i) Amount of heat absorbed at constant volume is internal energy change. (True)  
 (ii) The work done by the system is given the positive sign. (False)  
 (iii) Enthalpy is a state function but internal energy is not. (False)  
 (iv) Total heat content of a system is called enthalpy of the system. (True)  
 Q.13  $50\text{ cm}^3$  of 1.0 M HCl is mixed with  $50\text{ cm}^3$  of 1.00 M NaOH in a glass calorimeter. The temperature of the resultant mixture increases from  $21.0^\circ C$  to  $27.5^\circ C$ . Assume, that calorimeter losses of heat are negligible. Calculate the enthalpy change mole<sup>-1</sup> for the reactions. The density of solution to be considered is  $1\text{gcm}^3$  and specific heat is  $4.18\text{Jg}^{-1}\text{K}^{-1}$ . (H.W)

Ans: Volume of acid/or base used =  $V = 50\text{cm}^3$   
 Initial temperature =  $T_1 = 21^\circ C$   
 Final temperature =  $T_2 = 27.5^\circ C$   
 Rise in temperature =  $T = 27.5 - 21 = 6.5^\circ C$   
 Specific heat of water =  $s = 4.18\text{Jg}^{-1}\text{K}^{-1}$

Since density of water is around  $1\text{g cm}^{-3}$

Therefore mass of  $100\text{ cm}^3$  of solution =  $m = 100\text{g}$

Thus, heat evolved is given by

$$q = m \times s \times T = 100 \times 4.18 \times 6.5 = 2717\text{ J} = 2.717\text{ kJ}$$

$$\text{or } q = -2.717\text{ kJ (exothermic reaction)}$$

Since Molarity of acid/or base used =  $M = 1.0\text{ M}$  therefore

$1000\text{ cm}^3$  of solution contain HCl = 1.0 moles

$$50\text{ cm}^3 \text{ of solution contain HCl} = \frac{1}{1000} \times 50 = 0.05\text{ moles}$$

Hence

0.05 moles of HCl react with 0.05 moles of NaOH.



Thus heat of neutralization is given by  
0.05 moles of HCl produce heat

$$\frac{-2.717 \text{ kJ}}{0.05} = -54.34 \text{ kJ/mol}$$

1 moles of HCl produce heat

$$= -54.34 \text{ kJ/mol}$$

Hence heat of neutralization =  $\Delta H_n^\circ$

- Q.14 Hydrazine ( $\text{N}_2\text{H}_4$ ) is a rocket fuel. It burns in  $\text{O}_2$  give  $\text{N}_2$  and  $\text{H}_2\text{O}$ . 1.00 g of  $\text{N}_2\text{H}_4$  is burned in a bomb calorimeter. An increase of temperature  $3.51^\circ\text{C}$  is recorded. The specific heat of calorimeter is  $5.5 \text{ kJ K}^{-1}$ . Calculate the quantity of heat evolved. Also, calculate the heat of combustion of 1 mole of  $\text{N}_2\text{H}_4$ . (H.W)

$$\text{N}_2\text{H}_4 + \text{O}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$$

$$\begin{aligned} \text{Ans: Mass of hydrazine} &= m = 1 \text{ g} \\ \text{Heat capacity} &= c = 5.5 \text{ kJ K}^{-1} \\ \text{Rise in temperature} &= \Delta T = 3.51^\circ\text{C} \\ \text{Heat evolved during combustion is given by} \\ q &= c \times \Delta T = 5.5 \times 3.51 = 19.305 \text{ kJ} \\ \text{Since reaction is exothermic, therefore} \\ q &= -19.305 \text{ kJ} \end{aligned}$$

$$\text{Molecular mass of hydrazine (N}_2\text{H}_4) = 32 \text{ g mol}^{-1} = 1 \text{ mol}$$

$$\begin{aligned} \text{Since} \\ 1 \text{ g of hydrazine produce heat} &= -19.305 \text{ kJ} \\ 32 \text{ g of hydrazine produce heat} &= -19.305 \times 32 = -617.76 \text{ kJ/mole} \\ \text{Hence heat of combustion of hydrazine} &= \Delta H_c^\circ = -617.76 \text{ kJ/mole} \end{aligned}$$

- Q.15 Octane ( $\text{C}_8\text{H}_{18}$ ) is a motor fuel. 1.80 g of a sample of octane is burned in a bomb calorimeter having heat capacity  $11.66 \text{ kJ K}^{-1}$ . The temperature of the calorimeter increases from  $21.36^\circ\text{C}$  to  $28.78^\circ\text{C}$ . Calculate the heat of combustion for 1.8 g of octane. Also, calculate the heat for 1 mole of octane. (H.W)

$$\begin{aligned} \text{Ans: Mass of octane} &= m = 1.80 \text{ g} \\ \text{Heat capacity} &= c = 11.66 \text{ kJ K}^{-1} \\ \text{Initial temperature} &= T_1 = 21.36^\circ\text{C} \\ \text{Final temperature} &= T_2 = 28.78^\circ\text{C} \\ \text{Rise in temperature} &= \Delta T = 28.78^\circ\text{C} - 21.36 = 7.42^\circ\text{C} \end{aligned}$$

Heat evolved during combustion is given by

Heat evolved during combustion is given by

$$\begin{aligned} q &= c \times \Delta T = 11.66 \times 7.42 = 86.52 \text{ kJ} \\ \text{Since reaction is exothermic, therefore} \\ q &= -86.52 \text{ kJ} \end{aligned}$$

$$\text{Molecular mass of octane} = 114 \text{ g mol}^{-1} = 1 \text{ mol}$$

Since

$$1.80 \text{ g of octane produce heat} = -86.52 \text{ kJ}$$

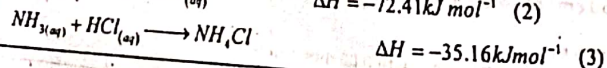
$$32 \text{ g of octane produce heat} = \frac{86.52}{1.8} \times 1 = -48.067 \text{ kJ}$$

$$114 \text{ g of octane produce heat} = \frac{86.52}{1.8} \times 114 = -5479.6 \text{ kJ/mole}$$

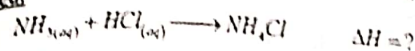
$$\text{Hence heat of combustion of octane} = \Delta H_c^\circ = -5479.6 \text{ kJ/mole}$$

- Q.16 By applying, Hess's law calculate the enthalpy change for the formation of an aqueous solution of  $\text{NH}_4\text{Cl}$  from  $\text{NH}_3$  gas and  $\text{HCl}$  gas. The results for the various reactions are as follows. (H.W)

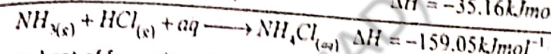
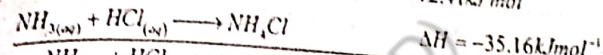
Ans: Given



Required:



Adding eq(1), (2) and (3)

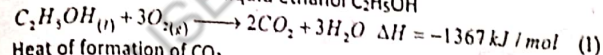


Hence heat of formation of aq solution on  $\text{NH}_4\text{Cl}$  is  $-159.05 \text{ kJ mol}^{-1}$

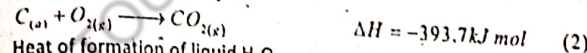
- Q.17 Calculate the heat of formation of ethyl alcohol from the following information  
Heat of combustion of ethyl alcohol is  $-1367 \text{ kJ mol}^{-1}$   
Heat of formation of carbon dioxide is  $-393.7 \text{ kJ mol}^{-1}$   
Heat of formation of water is  $-285.8 \text{ kJ mol}^{-1}$

Ans: Given

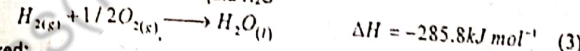
Heat of combustion of liquid ethanol  $\text{C}_2\text{H}_5\text{OH}$



Heat of formation of  $\text{CO}_2$

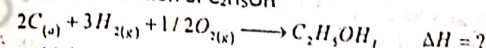


Heat of formation of liquid  $\text{H}_2\text{O}$



Required:

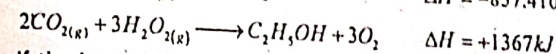
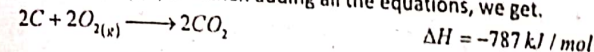
Heat of formation of  $\text{C}_2\text{H}_5\text{OH}$



Multiplying eq (2) by 2

Multiplying eq (3) by 3

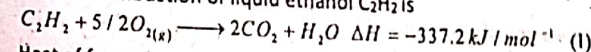
Reversing eq (1) and then adding all the equations, we get.



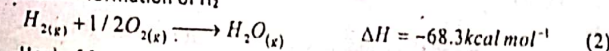
- Q.18 If the heats of combustion of  $\text{C}_2\text{H}_6$ ,  $\text{H}_2$  and  $\text{C}_2\text{H}_4$  are  $-337.2$ ,  $-68.3$  and  $-372.8 \text{ kJ mol}^{-1}$  respectively, then calculate the heat of the following reaction. (C.W)

Ans: Given

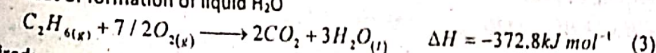
Heat of combustion of liquid ethanol  $\text{C}_2\text{H}_5\text{OH}$



Heat of formation of  $\text{H}_2$

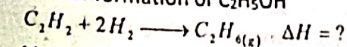


Heat of formation of liquid  $\text{H}_2\text{O}$



Required:

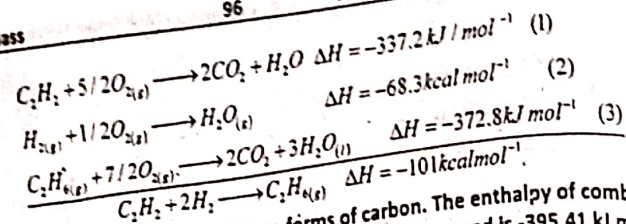
Heat of formation of  $\text{C}_2\text{H}_5\text{OH}$



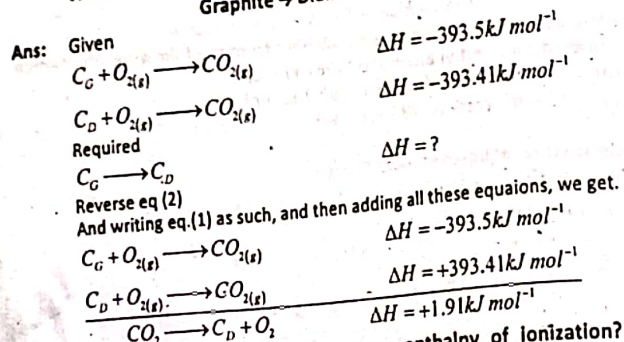
Multiplying eq (2) by 2

Reversing eq (1) and then adding all the equations, we get.

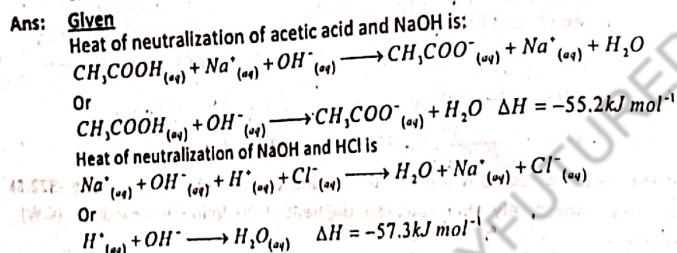
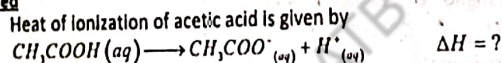




- Q.19 Graphite and diamond are two forms of carbon. The enthalpy of combustion of graphite at 25°C is -393.51 kJ mol<sup>-1</sup> and that of diamond is -395.41 kJ mol<sup>-1</sup>. What is the enthalpy change of the process?  
Graphite → Diamond at the same temperature? (H.W)

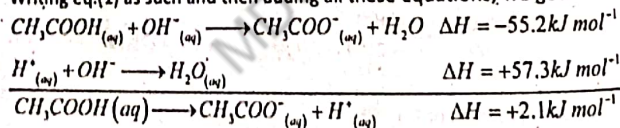


- Q.20 What is the meaning of the term enthalpy of ionization? If the heat of neutralization of HCl and NaOH is -57.3 kJ mol<sup>-1</sup> and heat of neutralization of CH<sub>3</sub>COOH with NaOH is -55.2 kJ mol<sup>-1</sup>, calculate the enthalpy of ionization of CH<sub>3</sub>COOH. (H.W)

**Required**

Reversing eq(2)

Writing eq.(1) as such and then adding all these equations, we get.



- Q.21 (a) Draw a complete, fully labelled Born Haber cycle for the determination of KBr.

Ans: Consult Text Book

## CHAPTER- 8 (CHEMICAL EQUILIBRIUM)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS-2020)

- The law of mass action was given by Guldberg and P. Waage in:
- (a) 1846 (b) 1946 (c) 1864 (d) 1909
  - The law of mass action was given by: (b) Bodenstein
  - The units of  $K_c$  for reaction  $\text{N}_2 + \text{O}_2 \rightleftharpoons 2\text{NO}$  will be: (c) Berthelot (d) Guldberg and Waage
  - The unit of equilibrium constant ( $K_c$ ) for the reaction  $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$  will be: (a) moles / dm<sup>3</sup> (b) mole<sup>-1</sup>/dm<sup>3</sup> (c) mole<sup>-2</sup>dm<sup>-6</sup> (d) No unit
  - Having no unit (b) Mol dm<sup>-3</sup> (c) Mol<sup>-2</sup> dm<sup>-6</sup> (d) Mol<sup>-2</sup> dm<sup>-6</sup> (3 times)
  - The relationship between  $K_p$  and  $K_c$  is given by: (b)  $K_c = K_p(P/N)^{\Delta n}$  (c)  $K_p = K_c(RT)^{\Delta n}$  (d)  $K_p = K_c(RT)^{-\Delta n}$  (3 times)
  - Acid having  $K_a > 1$  will be: (b) Very weak (c) moderate (d) strong
  - For the reaction  $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$  (a)  $K_c = K_p$  (b)  $K_c < K_p$  (c)  $K_c > K_p$  (d)  $K_c = K_p$
  - Reaction of  $\text{BiCl}_3$  with  $\text{H}_2\text{O}$  gives white ppt. of  $\text{BiOCl}$  and  $\text{HCl}$  is formed: The white ppt. disappears by (a) Adding  $\text{BiCl}_3$  (b) Adding  $\text{HCl}$  (c) Increasing temperature (d) Decreasing temperature
  - The value of  $K_w$  at 25°C is: (a)  $0.11 \times 10^{-14}$  (b)  $0.30 \times 10^{-14}$  (c)  $1 \times 10^{-14}$  (d)  $3 \times 10^{-14}$  (3 times)
  - The units for  $K_w$  of  $\text{H}_2\text{O}$  are: (a) mol dm<sup>-3</sup> (b) mol<sup>2</sup> dm<sup>-6</sup> (c) mol<sup>-2</sup> dm<sup>6</sup> (d) mol<sup>-2</sup> dm<sup>-3</sup> (3 times)
  - The optimum temperature for the synthesis of  $\text{NH}_3$  by Haber's process is: (a) 200°C (b) 300°C (c) 400°C (d) 500°C
  - Catalyst used in preparation of  $\text{NH}_3$  from  $\text{N}_2$  and  $\text{H}_2$  is: (a) Fe (b) Ni (c) Pt (d)  $\text{V}_2\text{O}_5$
  - In synthesis of ammonia by Haber's process, the optimum condition for pressure is: (a) 150-160 atm (b) 170-200 atm (c) 200-300 atm (d) 300-350 atm (1 times) (2017 = 1 time)
  - When  $\text{KCl}$  is added to a saturated solution of  $\text{KClO}_3$ , the equilibrium is shifted to the: (a) Forward Direction (b) Backward Direction (c) Not affected (d) All of these (2 times)
  - An aqueous solution of ethanol in water may have vapour pressure: (a) Equal to that of water (b) Equal to that of ethanol (c) More than that of water (d) Less than that of water
  - The increase in dilution of solution (a) Increases the degree of dissociation (b) Decreases the degree of dissociation (c) Does not affect the degree of dissociation (d) Depends upon the degree of the solute
  - A solution with  $\text{pH}=0$ , indicates molar concentration of  $\text{H}^+$  ions. (a)  $10^{-7}$  (b)  $10^{-14}$  (c)  $10^{-14}$  (d) 1.0
  - Molarity of pure water is: (a) 1 (b) 18 (c) 55.5 (d) 6 (4 times)
  - The nature of milk is: (a) Acidic (b) Basic (c) Neutral (d) Normal
  - The  $\text{pH}$  of  $10^{-4}$  moles/dm<sup>3</sup> of  $\text{Ba}(\text{OH})_2$  is: (a) 4.5 (b) 6.4 (c) 7.5 (d) 10.3
  - $\text{pH}$  of buffer can be calculated by using: (a) Moseley's equation (b) Henderson's equation (c) De-Broglie's equation (d) Bohr's equation
  - The  $\text{pH}$  of the gastric juice is: (a) 2.0 (b) 3.0 (c) 3.5 (d) 5.6



23. Which aqueous solution has highest pH?  
(a) 0.1M NaOH (b) 0.1M H<sub>2</sub>SO<sub>4</sub> (c) 0.1M HCl (d) 0.2M HNO<sub>3</sub>
24. The pH of buffer of CH<sub>3</sub>COONa and CH<sub>3</sub>COOH is:  
(a) 7 (b) >7 (c) <7 (d) 1
25. The term pH was introduced by:  
(a) Henderson (b) Sorenson (c) Goldstein (d) Thomson
26. The pH of human blood is:  
(a) 8.0 (b) 7.53 (c) 7.63 (d) 7.35
27. The solubility of KClO<sub>3</sub> in water is suppressed by adding:  
(a) NaClO<sub>3</sub> (b) NaCl (c) KMnO<sub>4</sub> (d) KCl
28. In the presence of common ion, the ionization of an electrolyte will:  
(a) Increase (b) decrease (c) no affect (d) Moderate change
29. The HCl is added to aqueous solution of H<sub>2</sub>S the solubility will be:  
(a) Increase (b) Remain constant (c) First decreases then increases (d) Decreases
30. When HCl is added to H<sub>2</sub>SO<sub>4</sub> aqueous solution, its ionization:  
(a) Increases (b) Decreases (c) First increases then decreases (d) Remain constant
31. The term pH was introduced by:  
(A) Henderson (B) Millikan (C) Le-chattilier (D) Sorenson
32. The pH of Milk of Magnesia is:  
(A) 10.5 (B) 3.5 (C) 8.5 (D) 11.1
33. For which system does the equilibrium constant, K<sub>c</sub> has units of (concentration)<sup>-1</sup>:  
(A) N<sub>2</sub> + 3H<sub>2</sub> ⇌ 2NH<sub>3</sub> (B) 2NO<sub>2</sub> ⇌ N<sub>2</sub>O<sub>4</sub> (C) H<sub>2</sub> + I<sub>2</sub> ⇌ 2HI (D) PCl<sub>5</sub> ⇌ PCl<sub>3</sub> + Cl<sub>2</sub>
34. Which one affects the value of K<sub>c</sub>?  
(A) concentration (B) temperature (C) pressure (D) catalyst
35. Which one of the following salt dissolves in water to form a solution with a pH greater than 7?  
(A) NaCl (B) CuSO<sub>4</sub> (C) Na<sub>2</sub>CO<sub>3</sub> (D) NH<sub>4</sub>Cl
36. Affects the value of K<sub>c</sub>.  
(A) concentration (B) temperature (C) catalyst (D) pressure
37. When ionic product of a solution is greater than the solubility product at particular temperature then the solution is said to be:  
(A) unsaturated (B) saturated (C) very dilute (D) super saturated
38. Approximate pH of apple is:  
(A) 2.7 (B) 3.1 (C) 4.2 (D) 4.5
39. The pH of a solution is 9, the solution is:  
(a) Weakly acidic (b) Weakly basic (c) Strongly acidic (d) strongly basic

## Answers

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
c	d	d	D	c	d	c	b	c	b	C	a	c	b	c
16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
a	D	c	a	d	b	a	a	c	b	d	d	b	c	C
31	32	33	34	35	36	37	38	39						
d	a	b	b	c	b	b	b	d						

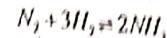
## CHAPTER - 8 (CHEMICAL EQUILIBRIUM)

### SHORT QUESTIONS

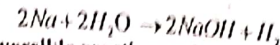
(ACCORDING TO ALP SMART SYLLABUS-2020)

1. Reversible and Irreversible Reactions:

1. Difference between reversible and irreversible reactions? (8 times)  
Ans: Reversible reaction: The reaction which can proceed both in forward as well as backward direction is called reversible reaction.



Irreversible reaction: The reaction which take place only in one direction is called irreversible reaction.



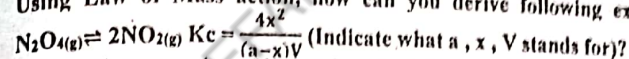
2. Explain the term reversible reaction and state of equilibrium?  
Ans: State of equilibrium: The state of reversible reaction in which rate of forward reaction becomes equal to rate of reverse reaction is called chemical equilibrium. It is always dynamic in nature.
3. Define State of Dynamic Equilibrium.  
Ans: When the forward and reverse reactions are taking place exactly at the same rate the equilibrium established is called dynamic equilibrium.

Law of Mass Action:

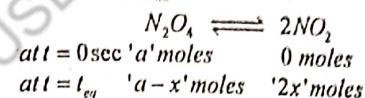
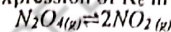
2. State law of mass action?

- Ans: Law of mass action: The rate at which substances react is directly proportional to their active masses. (8 times)

5. Using Law of Mass action, how can you derive following expression for



- Ans: N<sub>2</sub>O<sub>4</sub>(g), the expression of K<sub>c</sub> involves the factor of volume.



$$K_c = \frac{\frac{a-x}{V}}{\left[ \frac{NO_2}{N_2O_4} \right]^2} = \frac{\frac{2x}{V}}{\left[ \frac{a-x}{V} \right]^2}$$

$$K_c = \frac{4x^2}{(a-x)V}$$

'a' is the initial number of moles of N<sub>2</sub>O<sub>4</sub>, 'x' is number of moles of N<sub>2</sub>O<sub>4</sub> decomposed and 'V' is total volume of N<sub>2</sub>O<sub>4</sub> and NO<sub>2</sub> at equilibrium stage.

6. State law of Mass Action. Give equilibrium constant expression K<sub>c</sub> for the following reaction. (2 times)



- Ans: Law of mass action: The rate at which substance reacts is directly proportional to its active mass and the rate of reaction is directly proportional to the product of active masses of the reactants.



$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

3. Applications of Equilibrium Constant:

7. Write two applications of equilibrium constant? (2 times)

- Ans: There are following applications of equilibrium constant.  
(i) Prediction of direction of reaction. (ii) Extent of a chemical reaction.  
(iii) Effect of various factors on chemical equilibrium.
8. Why the equilibrium constant value has its units for some of the reversible reactions but has not units for some other reactions?

- Ans: Units of K<sub>c</sub> depend upon the number of moles of reactants and products involved in the reaction.  
(i) If number of moles of reactant and products are same it has no unit.



- (ii) If number of moles of reactants are different from products then  $K_c$  has units related to the concentration or pressure.  
 Reversible reaction attains the position of equilibrium which is dynamic in nature and not static. Explain it?  
 Ans: At equilibrium state the reaction is not stopped. Only the rate of forward reaction become equal to reverse reaction. Since reaction is in progress in both the directions. Therefore equilibrium is dynamics in nature not static.
10. The change of temperature disturbs both the equilibrium position and equilibrium constant of a reaction. Explain with reason?  
 Ans: According to Le-Chatelier's principle an increase in temperature will favour the endothermic reaction and decrease in temperature will favour the exothermic reaction. Therefore, change of temperature will disturb equilibrium position. The equilibrium constant is temperature dependent therefore with the change of temperature a new equilibrium position will be established. (5 times)
11. How  $K_c$  predict the extent of chemical reaction?  
 Ans: The value of  $K_c$  also helps us to predict extent of chemical reaction. There are three possibilities.  
 (i) Large  $K_c$  value: If  $K_c$  value is large it means reaction is almost completed.  
 (ii) Small  $K_c$  value: If  $K_c$  value is small it means reaction does not proceed appreciably in forward direction. Small amount of products will be formed.  
 (iii)  $K_c$  is in fractions: If  $K_c$  is in fractions. It means little forward reaction. (2 times)
12. Define  $K_f$  and  $K_b$  for reversible reactions?  
 Ans:  $K_f$  and  $K_b$ :  $K_f$  is rate constant for forward reaction and  $K_b$  is rate constant for reverse or backward reaction. At equilibrium, rate of forward reaction become equal to reverse reaction.
13. Give the physical significance of  $K_f$  and  $K_b$ ?  
 Ans: The rate of forward and reverse reaction tell us the condition on which a reaction will depend. It also tells about:  
 (i) Direction of a reaction.  
 (ii) Extent of a reaction.
14. Why do rates of forward reactions slow down when the reversible reaction approaches the equilibrium stage? (2 times)  
 Ans: The rate of forward reaction is directly proportional to molar conc. of reactants. Near the equilibrium stage, the concentration of reactant become small. Therefore the rate of forward reaction slow down due to decrease in concentration.
15. Derive value of  $K_c$  for the reaction  $\text{PCl}_5 \rightleftharpoons \text{PCl}_3 + \text{Cl}_2$ ?  
 Ans:  $\text{PCl}_5 \rightleftharpoons \text{PCl}_3 + \text{Cl}_2$

$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

16. Increasing pressure increase the oxidation of  $\text{SO}_2$  to  $\text{SO}_3$ . Explain why?



The number of moles of products are smaller than number of moles of reactants. Hence according to Le-Chatelier's principle, an increase in pressure will favour the forward reaction. During this process the pressure is kept at one atmosphere. Hence by increasing the pressure the conc. of oxygen increases, which results in greater yield of  $\text{SO}_3$ .

17. What are  $K_c$  and  $K_p$  and how these are related? (2 times 2018)

Ans:  $K_c$ : When the concentration of reactants or products are expressed in moles dm<sup>-3</sup> the equilibrium expression called  $K_c$ .  
 $K_p$ : When the reactants or products are gases then the conc. terms may be replaced by partial pressure called  $K_p$ .

$$K_p = K_c(RT)^{\Delta n}$$

Where  $R$  = Gas constant.

$\Delta n$  = Number of moles of products - number of moles of reactants.

### Le-Chatelier's Principle:

4. What will be the effect of volume change on the following system at equilibrium state?  
 (a)  $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$  (b)  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$  (3 times 2018)
- Ans:  $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ : The change in volume will affect the equilibrium position only. When volume is decreased reaction moves in backward direction to establish equilibrium again. But equilibrium constant is not affected.  
 $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ : When we decrease in volume for this reaction equilibrium position will shift in forward direction.
19. How does the change of pressure shifts the equilibrium position in the synthesis of ammonia?  
 $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$  (2 times)(2018)
- Ans: When we increase or decrease the pressure of a gaseous system the equilibrium position is disturbed. If moles of products and reactants are same in reaction no disturbance in equilibrium position take place. If moles of products and reactants are unequal, then reaction will move toward less number of moles by increasing pressure. In the formation of  $\text{NH}_3$  more product is formed by increasing pressure.
- $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$
20. What is the effect of catalyst of equilibrium constant? (6 times)  
 Ans: A catalyst does not change the equilibrium position. A catalyst only increases the rate of forward and reverse reaction and attain equilibrium earlier.
21. Why change of volume disturbs the equilibrium position for some of the gaseous phase reactions but not the equilibrium constant? (2 times)  
 $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$
- Ans: This gas phase reaction proceeds with the decrease in the number of moles and hence decrease in volume at equilibrium stage. When the reaction approaches the equilibrium stage, the volume of the equilibrium mixture is less than the volume of reactants taken initially. If one decrease the volume further at equilibrium state, the reaction is established a new equilibrium position while  $K_c$  remains constant. The reverse happens when the volume is increased or pressure is decreased at equilibrium stage.
22. A catalyst does not affect the equilibrium position and  $K_c$  of a reversible reaction. Explain?  
 Ans: In most of the reversible reactions, the equilibrium is not always reached within a suitable short time. So, an appropriate catalyst is added. A catalyst does not affect the equilibrium position of the reaction. It increases the rates of both forward and backward reactions and it reduces the time to attain the state of equilibrium. Actually a catalyst lowers the energy of activation of both forward and reverse steps by giving new path to the reaction.
23. State Le-Chatelier's principle? (4 times)
- Ans: Le-Chatelier's principle: If stress is applied to a system at equilibrium, the system will act in such a way so as to nullify as far as possible the effect of that stress.
24. Why during the synthesis of  $\text{NH}_3$  temperature is kept low?  
 $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$
- Ans: This is an exothermic reaction. Hence decrease in temperature will favour the forward reaction. That's why during  $\text{NH}_3$  preparation temperature kept low. Actually optimum temperature is 400 to 450°C.
25. What is the effect of increase in temperature on the yield of the product for the reaction.  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ ?  
 $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g}) + \text{Heat}$
- Ans: It is an exothermic reaction in the forward direction. It means that by increasing temperature, the reaction at equilibrium will shift the reaction to the backward direction by decreasing the yield of product.
26. What will be the effect of increase of pressure on the decomposition of  $\text{PCl}_5$ ?



Ans:

Since the number of moles of reactant are smaller than products. Therefore by increasing pressure equilibrium will shift towards left side and more amount of  $\text{PCl}_5$  will produce.

27. What are optimum conditions for the synthesis of  $\text{NH}_3$ ? (3 times)

Ans: At high pressure, low temperature and continuous removal of  $\text{NH}_3$  will give best yield of  $\text{NH}_3$ . The optimum conditions are:

- (i) 200-300 atm pressure.  
(ii) 400°C temperature  
(iii) Piece of iron as catalyst.
28. What will be the effect of increase of pressure and temperature on the following reactions? (3 times)

Ans:  $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$   
In this case number of moles of products are smaller than reactant. Therefore by increase in pressure will favour the forward reaction. Hence,  $\text{NH}_3$  concentration will increase. But as we know that it is an exothermic reaction. So, by increasing temperature yield of  $\text{NH}_3$  become lower.

29. What is the effect of temperature change on the direction of the reaction? (2 times)

Ans: According to Le-Chatelier's principle "Decreasing the temperature will shift it to the forward direction".

5. Ionic Product of Water: (2 times)

30. Define  $\text{pK}_a$  and  $\text{pK}_b$ ?  
Ans: The values of  $K_a$  and  $K_b$  for weak acids and bases are small numbers usually expressed in exponential form. It is convenient to convert them into whole numbers by taking their negative log. Thus we obtain  $\text{pK}_a$  and  $\text{pK}_b$  values of acids and bases.

$$\text{pK}_a = -\log K_a$$

$$\text{pK}_b = -\log K_b$$

31. Prove that  $\text{pK}_a + \text{pK}_b = 14$  at 25°C? (2 times)

Ans: For an acid  $K_a \times K_b = K_w$ , at 25°C the value of  $K_w$  is  $1 \times 10^{-14}$ . By taking the log of above equation we get

$$\log K_a + \log K_b = \log K_w$$

Multiply equation by -1, we get

$$-\log K_a + (-\log K_b) = -\log K_w$$

$$\text{pK}_a + \text{pK}_b = \text{pK}_w$$

$$\text{pK}_a + \text{pK}_b = 14 \quad (\text{pK}_w = -\log(1 \times 10^{-14}) = 14)$$

32. Define  $\text{pOH}$  of a solution. Give its equation? (2 times)

Ans: The negative logarithm of hydroxyl ion concentration is called  $\text{pOH}$  of that solution.

Mathematically it can be written as:

$$\text{pOH} = -\log[\text{OH}^-]$$

33. What is the justification for the increase of ionic product of water with temperature? (1 time)

Ans:  $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$   
Water is a very weak electrolyte. The ionization of  $\text{H}_2\text{O}$  is an endothermic process which is increased with the rise in temperature.

34. Is it true that value of  $K_w$  increases 75 times when temperature is increased from 0°C to 100°C? (2 times)

Ans: Ionic product of water is increase with rise in temperature because at high temperature ionization become easy. For example

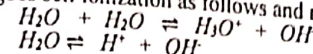
$$K_w \text{ at } 0^\circ\text{C} \text{ is } 0.10 \times 10^{-14}. \text{ At } 100^\circ\text{C } 7.5 \times 10^{-14}$$

Thus the value from 0°C-100°C is:

$$\frac{7.5 \times 10^{-14}}{0.10 \times 10^{-14}} = 75 \text{ times}$$

35. What is ionic product of water? How does this value change by change in temperature? (4 times)

Ans: Water undergoes self ionization as follows and reaction is reversible:



$$K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = 1.8 \times 10^{-16} \text{ moles.dm}^{-3}$$

Since the water is present in very large excess and very few of its molecules undergo ionization, so its concentration remains effectively constant. Constant concentration of water is taken on L.H.S. and multiplied with  $K_c$  to get another constant  $K_w$ .

$$1.8 \times 10^{-16} \times 55.5 = 1.01 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$$

This  $1.01 \times 10^{-14}$  is called  $K_w$  of water at 25°C

$$K_c[\text{H}_2\text{O}] = [\text{H}^+][\text{OH}^-]$$

$$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14} \text{ at } 25^\circ\text{C}$$

$K_w$  is called ionic product of water.

The value of  $K_w$  increases almost 75 times when temperature is increased from 0°C to 100°C. Anyhow, the increase in  $K_w$  is not regular.

36. Is it true that the sum of  $\text{pK}_a$  and  $\text{pK}_b$  is always equal to 14 at all temperature for any acid? (3 times)

Ans: It is not true because  $\text{pK}_a$  and  $\text{pK}_b$  values are temperature dependent. The degree of ionization of any acid increase as the temperature increases. Hence the value of  $\text{pK}_a$  and  $\text{pK}_b$  change with change in temperature.

37. What will be nature of solution when: (a)  $\text{pH}=3.0$  (b)  $\text{pH}=8.0$ ?

Ans: The value of  $\text{pH}$  varies between 0-14. A solution having  $\text{pH}$  value 0-7 are acidic in nature while a solution having  $\text{pH}$  value 7-14 are basic in nature.

(i)  $\text{pH} = 3$  This solution is acidic.

(ii)  $\text{pH} = 8$  This solution is basic.

38. Define  $\text{pH}$  and  $\text{pOH}$ . Give its equation? (10 times)

Ans:  $\text{pH}$ : The negative logarithm of  $\text{H}^+$  ions concentration is called  $\text{pH}$ .

$$\text{pH} = -\log[\text{H}^+]$$

$\text{pOH}$ : The negative logarithm of  $\text{OH}^-$  ions concentration is called  $\text{pOH}$ .

$$\text{pOH} = -\log[\text{OH}^-]$$

39. What will be nature of solution when (a)  $\text{pH}$  is more than 7 (b)  $\text{pH}$  is smaller than 7?

Ans:  $\text{pH}$  scale generally ranges from 0 to 14. When  $\text{pH}$  is less than 7 the solution has acidic nature. When  $\text{pH}$  is greater than 7 then solution is base.

40. Define  $\text{pH}$  and  $\text{pOH}$ . How are they related with  $\text{pK}_w$ ? (3 times)

$$\text{pK}_w = \text{pH} + \text{pOH}$$

Ans: Relation of  $\text{pH}$  and  $\text{pOH}$  with  $\text{pK}_w$ : We know that

$$K_w = [\text{H}^+][\text{OH}^-]$$

Taking -ve log on both sides

$$-\log K_w = (-\log[\text{H}^+]) + (-\log[\text{OH}^-])$$

Negative log of  $K_w$  is called  $\text{pK}_w$ . So,

$$\text{pK}_w = \text{pH} + \text{pOH}$$

41. Calculate  $\text{pH}$  of  $10^{-3} \text{ mol dm}^{-3} \text{ HCl}$  (3 times)

Ans:  $\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$

$$[\text{H}^+] = 10^{-3}$$

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log(10^{-3})$$

$$\text{pH} = -(-3)\log(10)$$

$$\text{pH} = 3 (1) = 3$$

42. Prove that  $\text{pH} + \text{pOH} = 14$ .



Ans: Water is a very poor conductor of electricity but its conduction is measurable. Self ionization of water is as under:



$$K_i = \frac{[H^+][OH^-]}{[H_2O]} = 1.8 \times 10^{-16} \text{ moles, dm}^{-3}$$

Since the water is present in very large excess and very few of its molecules undergo ionization, so its concentration remains effectively constant. Constant concentration of water is taken on L.H.S. and multiplied with  $K_i$  to get another constant  $K_w$ .

$$1.8 \times 10^{-16} \times 55.5 = 1.01 \times 10^{-14} = [H^+][OH^-]$$

This  $1.01 \times 10^{-14}$  is called  $K_w$  of water at  $25^\circ\text{C}$

$$K_a[H_2O] = [H^+][OH^-]$$

$$K_w = [H^+][OH^-] = 10^{-14} \text{ at } 25^\circ\text{C}$$

$$[H^+] = 10^{-7}$$

$$[OH^-] = 10^{-7}$$

$$pH = -\log[H^+]$$

$$pH = -\log(10^{-7}) = 7 \quad \dots\dots(i)$$

$$pOH = -\log[OH^-]$$

$$pOH = -\log(10^{-7}) = 7 \quad \dots\dots(ii)$$

Adding (i) and (ii).

$$pH + pOH = 7 + 7 = 14$$

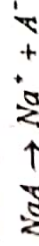
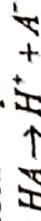
## 6. Ionization Constant of Acids & Bases:

43. Why aqueous solution of  $CH_3COONa$  is basic in nature? (2 times)

Ans: When salt  $CH_3COONa$  is dissolved in water, it produces  $NaOH$  and  $CH_3COOH$ .  $NaOH$  is a strong base while  $CH_3COOH$  is a weak acid. So, the solution becomes basic in nature.

44. Write down Henderson equation for acidic and basic buffers? (3 times)

Ans: Henderson's equation: Consider a reaction:



The ionization const. for acid.  $K_a = \frac{[H^+][A^-]}{[HA]}$

$$[H^+] = \frac{K_a[HA]}{[A^-]}$$

Taking -ve. log on both sides  $-\log[H^+] = -\log\left[\frac{K_a[HA]}{[A^-]}\right]$

$$-\log[H^+] = -\log K_a - \log\left[\frac{[HA]}{[A^-]}\right]$$

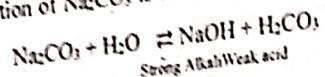
$$pH = pK_a - \log\left[\frac{[HA]}{[A^-]}\right]$$

$$pH = pK_a - \log\left[\frac{[acid]}{[salt]}\right]$$

$$pH = pK_a + \log\left[\frac{[salt]}{[acid]}\right]$$



Aqueous solution of  $\text{Na}_2\text{CO}_3$  is alkaline in nature as given by following chemical reaction.



7. **Buffer Solution:** (5 times)  
 Explain the terms buffer and buffer capacity?  
 51. **Ans:** Buffer: The solution which resist the change in pH by the addition of small amount of an acid or base is called buffer.  
 Buffer capacity: The capacity of a buffer to maintain definite pH is called buffer capacity.
52. **How does a buffer act? explain with an example?** (3 times)  
 Let us take the example of an acidic buffer consisting of  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COONa}$ . Common ion effect helps us to understand how the buffer will work.  $\text{CH}_3\text{COOH}$ , being a weak electrolyte undergoes very little dissociation.  

$$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$$

$$\text{CH}_3\text{COONa} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{Na}^+$$
 Actually a buffer mentioned above is a large reservoir of  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COO}^-$  components. When an acid or  $\text{H}_3\text{O}^+$  ions are added to this buffer, they will react with  $\text{CH}_3\text{COO}^-$  to give back acetic acid and hence the pH of the solution will almost remain unchanged.
53. **Write two uses of buffer solution?** (6 times)  
 53. **Ans:** (i) The pH of human blood is maintain at 7.35 with the help of buffers. A higher or lower value than this may prove fatal.  
 (ii) Buffers are used in many industrial processes like electroplating, dyes etc.
54. **Give preparation of acidic buffer solutions?** (3 times)  
 54. **Ans:** Acidic buffer solutions are prepared by mixing a weak acid and its salt with strong base. For example:  

$$\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$$

$$\text{CH}_3\text{COO}^- \text{Na}^+ \rightleftharpoons \text{CH}_3\text{COO}^- + \text{Na}^+$$

$$\text{CH}_3\text{COOH} / \text{CH}_3\text{COONa}$$
55. **A mixture of  $\text{NH}_4\text{OH}$  and  $\text{NH}_4\text{Cl}$  gives a basic buffer. Justify the statement?** (2 times)  
 55. **Ans:** A buffer which consists of a weak base and its salt with strong acid is called basic buffer. Since  $\text{NH}_4\text{OH}$  is a weak base and  $\text{NH}_4\text{Cl}$  is its salt with strong acid. Therefore it is basic buffer.
56. **Define buffer solution. Give an example of basic buffer?** (1 times)  
 56. **Ans:** Buffer solution: The solution that resists in pH changes when small amount of an acid or a base is added to it is called buffer solution. It is formed by mixing a weak base and its salt with strong acid. e.g.  $\text{NH}_4\text{OH} / \text{NH}_4\text{Cl}$
57. **Why we need buffer solution?** (2 times)  
 57. **Ans:** Buffers are the substances which resist the change in pH. Buffers are very important in many areas of chemistry. Buffers are needed in chemical analysis, pharmaceuticals, electroplating etc.
58. **What are acidic buffers? Explain with example.** (2 times)  
 58. **Ans:** A solution which resist a change in pH when small amount of acid is added to it is called acid buffer solution.  
 Preparation: These are prepared by mixing a weak acid and a salt of it with a strong base. e.g.  $\text{CH}_3\text{COOH} / \text{CH}_3\text{COONa}^+$
59. **What are buffer solutions? How a basic buffer can be prepared?**  
 59. **Ans:** Buffer solution: The solution that resists in pH changes when small amount of an acid or a base is added to it is called buffer solution.  
 Basic buffers are formed by mixing a weak base and its salt with strong acid. Such solution will give basic buffers with pH more than 7. Mixing of  $\text{NH}_4\text{OH}$  and  $\text{NH}_4\text{Cl}$  is one of the best example of such a basic buffer.

60. **What are Buffer Solutions? How Acidic Buffers are prepared?**  
 60. **Ans:** Those solutions, which resist the change in their pH when a small amount of an acid or a base is added to them, are called buffer solutions. They have specific constant value of pH and their pH values do not change on dilution and on keeping for a long time.  
 Acidic buffer is prepared by mixing a weak acid and a salt of it with a strong base. Such solutions give acidic buffers with pH less than 7. Mixture of acetic acid and sodium acetate is one of the best example of such a buffer.
61. **Differentiate between acidic and basic buffers.** (3 times)  
 61. **Ans:** The buffer which is prepared by mixing a weak acid and a salt of it with a strong base is known as acidic buffer. Such solutions give acidic buffers with pH less than 7. Mixture of acetic acid and sodium acetate is one of the best example of such a buffer.  
 The buffer which is prepared by mixing a weak base and a salt of it with a strong acid is known as basic buffer. Such solutions give basic buffers with pH more than 7. Mixture of  $\text{NH}_4\text{OH}$  and  $\text{NH}_4\text{Cl}$  is one of the best example of basic buffer.
62. **What are applications of buffer in daily life?**  
 62. **Ans:** Buffers are important in many areas of chemistry and allied sciences like molecular biology, microbiology, cell biology, soil sciences, nutrition and the clinical analysis.
7. **Common Ion Effect:**  
 63. **What is common ion effect? Give an example.** (6 times)  
 63. **Ans:** Common ion effect: The decrease in degree of ionization of a weak electrolyte by the addition of another electrolyte having a common ion is called common ion effect.  
 e.g.  $\text{NaCl} \rightleftharpoons \text{Na}^+ + \text{Cl}^-$   

$$\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$$
 $\text{Cl}^-$  is a common ion.
64. **Why solid  $\text{NH}_4\text{Cl}$  is added in qualitative analysis of 3<sup>rd</sup> group basic radical before adding  $\text{NH}_4\text{OH}$ ?**  
 64. **Ans:** 
$$\text{NH}_4\text{OH} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$$
 To provide low concentration of  $\text{OH}^-$  ion,  $\text{NH}_4\text{Cl}$  is added which is more soluble.  

$$\text{NH}_4\text{OH} \rightleftharpoons \text{NH}_4^+ + \text{Cl}^-$$
 $\text{NH}_4\text{Cl}$  decrease the  $\text{OH}^-$  ion conc. by suppressing ionization of  $\text{NH}_4\text{OH}$  due to common ion. That's why a mixture of  $\text{NH}_4\text{Cl} + \text{NH}_4\text{OH}$  is used as reagent for third group analysis.
65. **How NaCl can be purified by common ion effect?** (1 time)  
 65. **Ans:** The impurities like  $\text{CaCl}_2$ ,  $\text{MgCl}_2$ ,  $\text{Na}_2\text{SO}_4$  can be removed by the use of common ion effect.  $\text{HCl}$  gas is passed through saturated solution of  $\text{NaCl}$ .  

$$\text{NaCl} \rightleftharpoons \text{Na}_{(aq)}^+ + \text{Cl}_{(aq)}^-$$

$$\text{HCl} \rightleftharpoons \text{H}_{(aq)}^+ + \text{Cl}_{(aq)}^-$$
 Concentration of  $\text{Cl}^-$  ions increase by passing  $\text{HCl}$ . According to Le-Chatelier's principle the reaction will move in reverse direction to keep the value of  $K_c$  constant. Hence  $\text{NaCl}$  will precipitated out.
66. **What is the effect of common ion on solubility?** (4 times)  
 66. **Ans:** The solubility of a partially soluble electrolyte is decreased by the addition of a more soluble electrolyte having common ion. For example precipitation of  $\text{KClO}_3$ .  

$$\text{KClO}_3 \rightleftharpoons \text{K}^+ + \text{ClO}_3^-$$

$$\text{KCl} \rightleftharpoons \text{K}^+ + \text{Cl}^-$$
67. **What happens to the acidic and basic properties of aqueous solutions when pH varies from zero to 14?**  
 67. **Ans:** Properties of solutions when pH changes from 0 to 14.







- 11<sup>th</sup> Class
9. Calculate the pH (i)  $10^{-4}$  mol  $\text{dm}^{-3}$  of  $\text{HCl}$  only 50% dissociated ( $\text{H}_2\text{X}$  is dibasic acid)? (ii)  $1.0$  mol. $\text{dm}^{-3}$  of  $\text{H}_2\text{X}$ , which is only 50% dissociated ( $\text{H}_2\text{X}$  is dibasic acid)?
10. Calculate the pH of a buffer solution in which  $0.11$  molar  $\text{CH}_3\text{COONa}$  and  $0.09$  molar acetic acid solutions are present.  $K_a$  for  $\text{CH}_3\text{COOH}$  is  $1.85 \times 10^{-5}$ ? (6 times)
11. Benzoic acid  $\text{C}_6\text{H}_5\text{COOH}$  is a weak mono-basic acid ( $K_a = 6.4 \times 10^{-5}$  mol. $\text{dm}^{-3}$ ). What is the pH of a solution containing  $7.32$  g of sodium benzoate in one  $\text{dm}^{-3}$  of  $0.02$  mole  $\text{dm}^{-3}$  benzoic acid? (3 times)
12. Define pH and pOH. How are they related with  $\text{pK}_w$ ?
13. Prove that  $\text{pK}_a + \text{pK}_b = \text{pK}_w$ ?
14. What is ionic product of water? How does its value vary with change in temperature?
- Chapter # 8

**Solved Exercise Chapter # 8**  
**ALP SMART SYLLABUS 2020**

(C.W)

- Q1. Multiple choice questions  
For which system does the equilibrium constant,  $K_c$  has units of (concentration)<sup>-1</sup>
- i) (a)  $N_2 + 3H_2 \rightleftharpoons 2NH_3$  (b)  $H_2 + I_2 \rightleftharpoons 2HI$   
(c)  $2NO_2 \rightleftharpoons N_2O_4$  (d)  $2HF \rightleftharpoons H_2 + F_2$
- ii) Which statement about the following equilibrium is correct  
 $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$   $H = -188.3 \text{ kJ mol}^{-1}$   
(a) The value of  $K_c$  falls with a rise in temperature  
(b) The value of  $K_c$  falls with increasing pressure  
(c) Adding  $V_2O_5$  catalyst increase the equilibrium yield of sulphur trioxide  
(d) The value of  $K_c$  is equal to  $K_p$
- iii) The pH of  $10^{-3} \text{ mol dm}^{-3}$  of an aqueous solution of  $H_2SO_4$  is  
(a) 3.0 (b) 2.7 (c) 2.0 (d) 1.5
- iv) An excess of aqueous silver nitrate is added to aqueous barium chloride and precipitate is removed by filtration. What are the main ions in the filtrate?  
(a)  $Ag^+$  and  $NO_3^-$  only (b)  $Ag^+$  and  $Ba^{2+}$  and  $NO_3^-$   
(c)  $Ba^{2+}$  and  $NO_3^-$  only (d)  $Ba^{2+}$  and  $NO_3^-$  and  $Cl^-$  (C.W)
- Q2. Fill in the blanks  
Law of mass action states that the rate at which a reaction proceeds, is directly proportional to the product of the active masses of the reactant.  
In an exothermic reversible reaction, decreases in temperature will shift the equilibrium towards the forward direction.  
The equilibrium constant for the reaction  $2O_3 \rightleftharpoons 3O_2$  is  $10^{55}$  at  $25^\circ\text{C}$ , it tells that ozone is unstable at room temperature.  
In a gas phase reaction, if the number of moles of reactants are equal to the number of moles of the products,  $K_c$  of the reaction is equal to the  $K_p$ .  
Buffer solution is prepared by mixing together a weak base and its salt with strong acid. (C.W)
- Q3. Label the sentences as True or False.  
i) When a reversible reaction attains equilibrium both reactants and products are present in a reaction mixture. (True)  
ii) The  $K_c$  of the reaction (False)

$A + B \rightleftharpoons C + D$   
is given by

$$K_c = \frac{[C][D]}{[A][B]}$$

therefore it is assumed that  $[A] = [B] = [C] = [D]$

- iii) A catalyst is a substance which increases the speed of the reaction and consequently increases the yield of the product. (False)
- iv) expression  $K_w = [H^+][OH^-] = 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$  and is represented by an (True)
- Q6 When a graph is plotted between time on x-axis and the concentrations of reactants and products on y-axis for a reversible reaction, the curves become parallel to time axis at a certain stage. (H.W.)
- (a) At what stage the curves become parallel?
- Ans: When the graph for a reversible reaction becomes parallel to x-axis, it means there is no further increase or decrease of reactants and/or products. This is the stage of dynamic equilibrium.
- (b) Before the curves become parallel, the steepness of curves falls! Give reasons.
- Ans: The curve for products rises very sharply in the beginning because conc. is maximum in the beginning. Since conc. of reactant falls with time so curve shows slow rate of fall in the graph paper. In case of products graph there sharp rise in the beginning because reactant concentration are more. A stage comes when it becomes parallel to time axis.
- (c) The rate of decrease of concentrations of any of the reactants and rate of increase of concentrations of any of the products may or may not be equal, for various types of reactions, before the equilibrium time. Explain it.
- Ans: When the number of moles of reactants and products are equal, then the rates of increase and decrease of concentration are equal. But when the number of moles in the balance chemical equation are different, Then the rate of increase of concentration of species will be different from the others.
- The dissociation of  $N_2O_4$  to give  $NO_2$  is reversible reaction. One mole of  $N_2O_4$  breaks and two moles of  $NO_2$  are generated. The speed with which the graph of  $NO_2$  rises is greater than the speed of falling of the graph of  $N_2O_4$ .



- Q7 (a) Write down the relationship of different types of equilibrium constants i.e.  $K_c$  and  $K_p$  for the following general reaction. (H.W)



Ans:  $K_C = \frac{C_C^c C_D^d}{C_A^a C_B^b}; K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}; K_X = \frac{X_C^c X_D^d}{X_A^a X_B^b}; K_n = \frac{n_C^c n_D^d}{n_A^a n_B^b}$

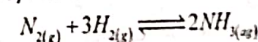
$$K_p = K_c (RT)^{\Delta n}$$

$$K_p = K_X(P)^{\Delta_T}$$

$$K_p = K_n \left( \frac{P}{N} \right)^{\Delta n}$$

- (b) Decide the comparative magnitudes of  $K_c$  and  $K_p$  for the following reversible reactions.

- Ans: (i) The ammonia synthesis is



$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_p = \frac{PNH_3}{PN_2 \cdot P^3H_2}$$

Since  $K_p = K_c (RT)^{\Delta n}$



$\Delta n$  = number of moles of products - no. of moles of reactants

$\Delta n = 2 - 4 = -2$

Now there are two situations.

(1)  $R = 0.0821$ , now when T is having such as value of  $RT < 1$  Then  $K_p < K_c$

(2)  $R = 0.0821$ , now when T is having such as value of  $RT > 1$  Then  $K_p > K_c$

$K_p = K_c (P)^{-2} = K_c \left(\frac{1}{P^2}\right)$

here again the value of P determine about relative smallness of  $K_p$  and  $K_c$ .

(ii)  $\Delta n = +1$

$PCl_5 \rightleftharpoons PCl_3 + Cl_2$

The discussion of  $NH_3$  synthesis of  $NH_3$ .

$K_p = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{P[PCl_3]P[Cl_2]}{P[PCl_5]}$

Since  $K_p = K_c (RT)^{\Delta n}$

$\Delta n = 2 - 1 = 1$

Now there are two situations.

(1)  $R = 0.0821$ , now when T is having such as value of  $RT < 1$  Then  $K_p < K_c$

(2)  $R = 0.0821$ , now when T is having such as value of  $RT > 1$  Then  $K_p > K_c$

$K_p = K_c (P)^{-1}$

When P is more than one atmosphere then  $K_p < K_c$ .

When P is less than one atmosphere then  $K_p > K_c$ .

$K_p = K_c \left(\frac{P}{N}\right)^{-1}$

$K_p = K_c \left(\frac{P}{N}\right)^{-1}$

Explain the following with reasons.

(a) The change of volume disturbs the equilibrium position for some of the gaseous phase reactions but not the equilibrium constant.

(b) The change of temperature disturbs both the equilibrium position and the equilibrium constant of a reaction.

Ans: (a) Those gaseous phase reversible reactions which happen with changing number of moles are affected by the change of volume at equilibrium stage. Their equilibrium position is distributed but equilibrium constant is not changed.

(b) All the reversible reactions are distributed by changing their equilibrium position and equilibrium constant by disturbing the temperature. Actually, change of temperatures also changes the equilibrium position. However, forward and reverse reactions are such that they have equal exothermicity and endothermicity then temperature effect will not disturb the  $K_c$  value.

Q11(a) What is an ionic product of water? How does this value vary with the change in temperature? Is it true that its value increase 75 times when the temperature of water is increased from 0°C to 100°C.

(b) What is the justification for the increase of ionic product with temperature?

(c) How would you prove that at 25°C, 1 dm<sup>3</sup> of water contains 10<sup>-7</sup> moles of H<sup>+</sup> and 10<sup>-7</sup> moles of OH<sup>-</sup>

Ans Since dissociation of H<sub>2</sub>O is endothermic process. So more H<sup>+</sup> and OH<sup>-</sup> are produced at elevated temperature. However, the conc. Of H<sup>+</sup> and OH<sup>-</sup> remains equal.

(b) Since  $[H^+][OH^-] = 10^{-14}$  at 25°C, so each  $[H^+] = 10^{-7}$  mol dm<sup>-3</sup> and  $[OH^-] = 10^{-7}$  mol dm<sup>-3</sup>

(c) What are buffer solutions? Why do we need them in daily life? (H.W)

Ans: Consult Text book Page No. 237

(b) How does the mixture of sodium acetate and acetic acid give us the acidic buffer?

Ans: Consult Text book Page No. 238

(c) Explain that a mixture of NH<sub>4</sub>OH and NH<sub>4</sub>Cl gives us the basic buffer.

(d) The mixture of NH<sub>4</sub>OH and NH<sub>4</sub>Cl gives a buffer solution. The equations are:

$NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$

Since OH<sup>-</sup> ions are produced in the solution, pH will be more than 7, and solution will be basic in character.

How do you justify that the greater quantity of CH<sub>3</sub>COONa in acetic acid decreases the dissociating power of acetic acid and so the pH increases.

(d)  $CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$

Ans:  $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$CH_3COO^- Na^+$  is a strong electrolyte. It suppresses the ionization of CH<sub>3</sub>COOH by common ion effect. For this, there is low conc. Of H<sup>+</sup> and so pH increases.

K<sub>c</sub> value for the following reaction is 0.016 at 520°C

Q19  $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$

Equilibrium mixture contains [HI] = 0.08 M, [H<sub>2</sub>] = 0.01 M, [I<sub>2</sub>] = 0.01 M. To this mixture more HI is added so that its new concentration is 0.096 M. What will be the concentration of [HI], [H<sub>2</sub>] and [I<sub>2</sub>] when equilibrium is re-established.

Ans:  $K_c = 0.016$

Chemical equation:  $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$

Initial conc. In mole dm<sup>-3</sup>:  $0.08 \rightleftharpoons 0.01 + 0.01$  = equilibrium

On adding more HI  $0.096 \rightleftharpoons 0.01 + 0.01$

Change in conc. dm<sup>-3</sup>:  $-2x \rightleftharpoons x + x$

Equilibrium conc. In moles/dm<sup>3</sup>  $(0.096 - 2x) \rightleftharpoons (0.01 - x) + (0.01 + x)$

Concentration of HI, H<sub>2</sub> and I<sub>2</sub> at new equilibrium = ?

Formula applied: The equilibrium expression for the reaction can be written as:

$K_c = \frac{[H_2][I_2]}{[HI]^2}$

On substituting the values, we get.

$0.016 = \frac{(0.01 + x) \times (0.01 + x)}{(0.096 - 2x)^2} = \frac{(0.01 + x)^2}{(0.096 - 2x)^2}$

On taking the square root on both sides.

$\sqrt{0.016} = \sqrt{\frac{(0.01 + x)^2}{(0.096 - 2x)^2}}$



$$0.126 = \frac{0.01+x}{0.096-2x}$$

$$0.126(0.096-2x) = 0.01+x$$

$$0.0121 - 0.252x = 0.01+x$$

$$1.252x = 0.0021$$

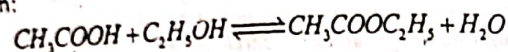
$$x = \frac{0.0021}{1.252} = 0.00168$$

Hence the equilibrium conc. of  $H_2 = (0.01 + 0.00168) = 0.01168M$   
 $I_2 = (0.01 + 0.00168) = 0.01168M$   
 $HI = (0.096 - 2 \times 0.00168) = 0.0826M$

Q20 The equilibrium constant for the reaction between acetic acid and ethyl alcohol is 4.0. A mixture of 3 moles of acetic acid and one mole of  $C_2H_5OH$  is allowed to come to equilibrium. Calculate the amount of ethyl acetate at equilibrium stage in number of moles and grams. Also calculate the masses of reactants left behind. (C.W)

Ans: Data: Equilibrium constant  $K_c = 4$   
 Initial conc. of acetic acid,  $(CH_3COOH) = 3 \text{ mole}$   
 Initial conc. of ethyl alcohol  $C_2H_5OH = 1 \text{ mole}$   
 Amount of ethyl acetate present at equilibrium = ?  
 Let the number of moles of ethyl acetate at equilibrium =  $x$  moles

Chemical equation:



Initial conc. In moles:  $3 + 1 \rightleftharpoons 0 + 0 \quad t=0$

Equilibrium conc. In moles  $dm^{-3}$



Formula applied.

Equilibrium constant expression for the reaction can be written as:

$$K_c = \frac{[CH_3COOC_2H_5][H_2O]}{[CH_3COOH][C_2H_5OH]}$$

Substituting the values.

$$4 = \frac{(x) \times (x)}{(3-x) \times (1-x)} = \frac{x^2}{x^2 - 4x + 3}$$

By cross multiplying

$$4(x^2 - 4x + 3) = x^2$$

$$4x^2 - 16x + 12 = x^2$$

$$4x^2 - x^2 - 16x + 12 = 0$$

$$3x^2 - 16x + 12 = 0$$

Formula applied

To find out the value of  $x$ , following formula is applied.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Here  $a = 3$ ,  $b = -16$  and  $c = 12$

Putting these values

$$x = \frac{+16 \pm \sqrt{(-16)^2 - 4(3)(12)}}{2(3)}$$

$$x = \frac{+16 \pm \sqrt{256 - 144}}{6}$$

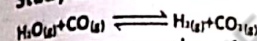
$$x = \frac{+16 \pm \sqrt{112}}{6}$$

$$x = 4.43 \text{ or } 0.9 \text{ mole}$$

The value,  $x = 4.43$  moles is not possible as it is greater than the concentrations of the reactants. Thus 0.9 mole of ethyl acetate is present at equilibrium.

Number of moles of ethyl acetate at equilibrium = 0.9 moles  
 Mass of ethyl acetate =  $0.9 \times 88 = 79.46g$   
 Number of moles of water produced = 0.9 moles  
 Mass of water produced =  $0.9 \times 18 = 16.2g$   
 Moles of acetic acid left behind =  $3.0 - 0.9 = 2.1 \text{ moles}$   
 Mass of acetic acid left behind =  $2.1 \times 60 = 126g$   
 Moles of alcohol left behind =  $1.0 - 0.9 = 0.1$   
 Mass of alcohol left behind =  $0.1 \times 46 = 4.6g$   
 Study the equilibrium (C.W)

Q21

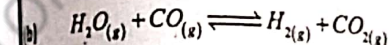


Write an expression of  $K_c$ .

(a) When 1.00 mole of steam and 1.00 mole of carbon monoxide are allowed to reach equilibrium, 33.3 % of the equilibrium mixture is hydrogen. Calculate the value of  $K_c$ . State the units of  $K_c$ .

(b) This is reversible reaction and all the reactants and products are in the gaseous state. It is homogeneous equilibria. So the value of  $K_p$  can be written in terms of partial pressures of gases at equilibrium stage.

$$K_p = \frac{P_{H_2} \cdot P_{CO_2}}{P_{H_2O} \cdot P_{CO}}$$



$$1.00 + 1.00 \rightleftharpoons 0 + 0 \quad \text{time} = 0 \text{ sec}$$

$$(1-x) + (1-x) \rightleftharpoons x + x \quad \text{time} = \text{equilibrium}$$

Total number of moles of reactants and products at equilibrium stage is 2. The number of moles of  $H_2$  i.e. 'x' is 33.3% of the mixture.

$$\frac{\text{number of moles of } H_2}{\text{total moles of mixture}} = \frac{33.3}{100}$$

$$\text{So, } \frac{x}{2} = \frac{33.3}{100} = \frac{6.6}{100} = 0.666$$

Hence at equilibrium stage the number of moles of hydrogen is 0.666.

Equilibrium concentrations:

$$[H_2O] = 1 - 0.66 = 0.33$$

$$[CO] = 1 - 0.66 = 0.33$$

$$[H_2] = 0.66$$

$$[CO_2] = 0.66$$

Formula applied.

$$K_p = \frac{[H_2][CO_2]}{[H_2O][CO]}$$

$$\text{Putting the values } K_p = \frac{0.66 \times 0.66}{0.33 \times 0.33} = 4$$

(C.W)

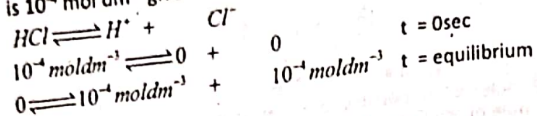
Calculate the pH of  $10^{-4} \text{ mole } dm^{-3}$  of HCl (b)  $10^{-4} \text{ mole } dm^{-3}$  of  $Ba(OH)_2$

$1.0 \text{ mole } dm^{-3}$  of  $H_2X$ , which is only 50% dissociated.

$1.0 \text{ mole } dm^{-3}$  of  $NH_4OH$  which is 1% dissociated.



Ans(a) HCl is strong electrolyte. It is 100% dissociated. So the concentration of HCl which is  $10^{-4} \text{ mol dm}^{-3}$  give  $\text{H}^+$  having concentration  $10^{-4} \text{ moles dm}^{-3}$ .



Data:

$$[\text{H}^+] = 10^{-4}$$

$$\text{pH} = ?$$

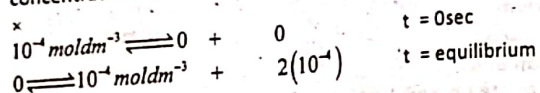
Formula applied

$$\text{pH} = -\log [\text{H}^+]$$

Putting the values.

$$\text{pH} = -\log 10^{-4} = +4 \log 10 = 4$$

(b)  $\text{Ba}(\text{OH})_2$  is also strong electrolyte and is 100% dissociated at very low concentration of  $10^{-4} \text{ mol dm}^{-3}$



$$[\text{OH}^-] = 2(10^{-4}) = 2 \times 10^{-4}$$

Formula applied

$$\text{pOH} = -\log [\text{OH}^-]$$

Putting the values.

$$\text{pOH} = -\log 2 \times 10^{-4} = 3.69$$

$$\text{Since } \text{pH} + \text{pOH} = 14$$

$$\text{Therefore } \text{pH} = 14 - \text{pOH} = 10.31$$

(c)  $\text{H}_2\text{X}$  is dibasic acid. One molecule of  $\text{H}_2\text{X}$  gives two ions of  $\text{H}^+$ . So one mole of  $\text{H}_2\text{X}$  will give 2 moles of  $\text{H}^+$ . But the acid is weak and it is 50% dissociated. So the concentration of  $\text{H}^+$  is  $1 \text{ mole dm}^{-3}$ .

$$[\text{H}^+] = 1 \text{ mol dm}^{-3}$$

$$\text{pH} = ?$$

Formula applied:

$$\text{pH} = -\log [\text{H}^+]$$

Putting the values

$$\text{pH} = -\log 1 = 0$$

(d)  $\text{NH}_4\text{OH}$  is a weak electrolyte and is 1% dissociated only. From one mole  $\text{dm}^{-3}$  only 1% is dissociated.

Data:

$$[\text{OH}^-] = 1 \times \frac{1}{100} = 0.01 = 10^{-2} \text{ mol dm}^{-3}$$

$$\text{pH} = ?$$

Formula applied.

$$\text{pOH} = -\log [\text{OH}^-] = -\log 10^{-2} = 2$$

$$\text{Since } \text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH} = 14 - 2 = 12$$

Q23(a) Benzoic acid,  $\text{C}_6\text{H}_5\text{COOH}$ , is a weak mono-basic acid ( $K_a = 6.4 \times 10^{-5} \text{ mol dm}^{-3}$ ). What is the pH of a solution containing 7.2 g of sodium benzoate in one  $\text{dm}^3$  of 0.02 mole  $\text{dm}^{-3}$  benzoic acid. (C.W)

Ans: Benzoic acid and sodium benzoate give a buffer solution.

$$\text{Mass of sodium benzoate} = 7.2 \text{ g dm}^{-3}$$

$$\text{Molar mass of sodium benzoate} = 144 \text{ g mol}^{-1}$$

$$\text{Number of moles of sodium benzoate} = \frac{7.2}{144} = 0.05 \text{ dm}^{-3}$$

$$K_a \text{ of benzoic acid} = 6.4 \times 10^{-5}$$

$$\text{p}K_a \text{ of benzoic acid, } -\log (6.4 \times 10^{-5}) = 4.2$$

Formula applied:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{sodium benzoate}]}{[\text{Benzoic acid}]}$$

Substituting the values.

$$\text{pH} = 4.2 + \log \frac{(0.05)}{(0.02)} = 4.2 + 0.39 = 4.59$$

(b) A buffer solution has been prepared by mixing 0.2 M  $\text{CH}_3\text{COONa}$  and 0.5 M  $\text{CH}_3\text{COOH}$  in 1  $\text{dm}^3$  of solution. Calculate the pH of solution.  $\text{p}K_a$  of acid = 4.74 at 25°C. How the values of pH will change by adding 0.1 mole of NaOH and 0.1 mole of HCl separately.

Data

$$[\text{CH}_3\text{COOH}] = 0.5 \text{ M}$$

$$\text{p}K_a = 4.74$$

Formula applied.

$$[\text{CH}_3\text{COONa}] = 0.2 \text{ M}$$

$$\text{pH} = ?$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]}$$

Substituting the values of  $\text{p}K_a$ ,  $\text{CH}_3\text{COONa}$  and  $\text{CH}_3\text{COOH}$ , we get

$$\text{pH} = 4.74 + \log \frac{[0.2 \text{ M}]}{[0.5 \text{ M}]} = 4.74 + \log (0.4) = 4.74 + (-0.4)$$

$$\text{pH} = 4.74 - 0.4 = 4.34$$

Addition of 0.1 Moles of NaOH:

Since NaOH is a strong acid therefore it is 100% dissociated. It generates 0.1 moles of  $\text{OH}^-$  ions. Out of 0.5 moles of  $\text{CH}_3\text{COOH}$ , 0.1 moles will react with  $\text{OH}^-$  ions and 0.4 moles of  $\text{CH}_3\text{COOH}$  is left behind in one  $\text{dm}^3$  of solution. This neutralization makes identical change in the amount of  $\text{CH}_3\text{COONa}$  and its concentration will also increase from 0.2 moles to 0.3 moles. Now putting these new values in Henderson equation, we get new value of pH.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]}$$

Substituting the values of  $\text{p}K_a$ ,  $\text{CH}_3\text{COONa}$  and  $\text{CH}_3\text{COOH}$ , we get

$$\text{pH} = 4.74 + \log \frac{[0.3]}{[0.4]} = 4.74 + \log (0.75) = 4.74 - 0.12 = 4.62$$

Addition of 0.1 Moles of HCl:

Since HCl is a strong base therefore it is 100% dissociated. It generates 0.1 moles of  $\text{Cl}^-$  ions. Out of 0.2 moles of  $\text{CH}_3\text{COONa}$ , 0.1 moles will react with  $\text{Cl}^-$  ions and 0.1 moles of  $\text{CH}_3\text{COONa}$  is left behind in one  $\text{dm}^3$  of solution. This neutralization makes identical change in the amount of  $\text{CH}_3\text{COOH}$  and its concentration will also increase from 0.5 moles to 0.6 moles. Now putting these new values in Henderson equation, we get new value of pH.

$$\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]}$$

Substituting the values of  $\text{p}K_a$ ,  $\text{CH}_3\text{COONa}$  and  $\text{CH}_3\text{COOH}$ , we get

$$\text{pH} = 4.74 + \log \frac{[0.1]}{[0.6]} = 4.74 + \log (-0.78) = 4.74 - 0.78 = 3.96$$

New value of pH shows that a minor change has taken place. On addition of NaOH its basic strength increases slightly but when HCl is added to the solution its acidity strength increases slightly.



## CHAPTER -9 (SOLUTIONS) OBJECTIVES (MCQ'S) (ACCORDING TO ALP SMART SYLLABUS-2020)

- Which one is not equation of Raoult's law?  
(a)  $\Delta P = P^\circ X_2$  (b)  $PV = n_2 RT$  (c)  $\frac{\Delta P}{P^\circ} = X_2$  (d)  $P = P^\circ X_1$
- Which one is not an electrolyte?  
(a) Aqueous NaCl (b) Aqueous  $\text{CuSO}_4$  (c) Cu metal (d)  $\text{H}_2\text{SO}_4$
- Relative lowering of vapour pressure is equal to:  
(a) Mole fraction of solute (b) Mole fraction of solvent (c) Molality (d) Molarity
- Melting point of ice can be lowered by the use of:  
(a) LiCl (b)  $\text{BeCl}_2$  (c) NaCl (d) AgCl
- A thermometer used in Landsberger's method can read up to:  
(a) 0.1K (b) 0.01F (c) 0.01K (d) 0.01°C
- Elevation of boiling point is:  
(a) Additive property (b) Constitutive property (c) Substitution property (d) Colligative property

**2018**

- Which one is used as antifreeze in radiator of automobile?  
(A) Aspartame (B) Ethylene glycol (C) Serotonin (D) Hydrazine
- Ideal solutions obey:  
(A) Henry's law (B) Avogadro's law (C) Raoult's law (D) Smith's law
- An aqueous solution of ethanol in water may have vapour pressure: (2 times)  
(A) equal to water (B) more than that of water (C) equal to ethanol (D) less than that of water

**Answers**

1	2	3	4	5	6	7	8	9
b	c	a	c	c	c	b	c	b

## CHAPTER-9 (SOLUTIONS) SHORT QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

- One molal solution of glucose in water is dilute as compared to one molar solution of glucose. Justify? (4 times)  
Ans: Both one molar and one molal solution of glucose contain one mole of substance. we can say equal number of particles i.e.  $6.02 \times 10^{23}$ . For one molar solution 1 mole of glucose is dissolved in 1 kg of  $\text{H}_2\text{O}$ . For one molal solution, 1 mole of glucose is taken in a flask and then made a volume of 1 dm<sup>3</sup>. Hence 1 molar solution contains less amount of  $\text{H}_2\text{O}$  than 1 molal solution.
- One molal solution of urea is more dilute as compared to its one molar solution. Why? (6 times)  
Ans: One molal solution contains one mole of urea per kg of solvent, while one molar solution contains one mole of urea per one dm<sup>3</sup> of solution. In one molal solution amount of water is greater than one molar solution because amount of solute is included in it.
- Why concentration in term of molality is independent of temperature while molarity depends upon temperature? (4 times)

- Ans: In molality, mass of solvent is taken. Mass is independent of temperature. In molarity volume of solution is taken. Volume changes by changing the temperature. So, concentration in term of molality is independent of temperature while molarity depends upon temperature.
- One molal solution of urea in water is dilute as compared to one molar solution of urea but the number of particles of solute is same. Justify it? (2 times)  
Ans: Both one molar and one molal urea solution contains equal number of molecules ( $6.022 \times 10^{23}$ ). Number of water molecules are different in both cases. One molal solution contains 1000g or 55.5 moles of water. One molar solution has less number of water molecules than one molal.
  - The sum of all the mole fractions is equal to one (unity). Discuss? (3 times)  
Ans: Consider a solution which consists of two components A and B. Their mole fraction  $X_A$  and  $X_B$  are given as:

$$X_A = \frac{n_A}{n_A + n_B}, \quad X_B = \frac{n_B}{n_A + n_B}$$

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B}$$

$$X_A + X_B = \frac{n_A + n_B}{n_A + n_B} = 1$$

Hence it is proved that sum of mole fraction for only solution is always unity.

- The total volume of the solution by mixing 100 cm<sup>3</sup> of water with 100 cm<sup>3</sup> of alcohol may not be equal to 200 cm<sup>3</sup> justify it?  
Ans: Alcohol and water are mix in all proportions. However, the properties of such solutions are not strictly additive. Generally, the volume decreases on mixing but in some cases it increases. Heat may be evolved or absorbed during the formation of such solutions.

- Raoult's Law:**  
Give two definitions of Raoult's law? (4 times)

- Raoult's law: (i). The vapour pressure of a solvent above a solution is equal to the product of the vapour pressure of pure solvent and mole fraction of solvent in solution.

$$P = P^\circ X_1$$

- (ii). Relative lowering of vapour pressure of a solution is equal to the mole fraction of solute.

$$\frac{\Delta P}{P} = X_2$$

- Why the non-ideal solutions do not obey the Raoult's law? (3 times)  
Ans: Many solutions do not behave ideally as they show deviations from Raoult's law due to difference in their molecular structure i.e. size, shape and intermolecular forces. During their formation change in volume and enthalpy take place. Hence they show deviation from ideality.
- Differentiate between ideals and non-ideal solutions? (9 times)

Ans:

Ideal solution	Non-ideal solution
(i) The solutions which obey Raoult's law are called ideal solutions.	The solutions which do not obey Raoult's law are called non-ideal solution.
(ii) In these solution, enthalpy change is zero.	In these solutions, enthalpy change is not zero. $\Delta H \neq 0$ e.g. Ethanol-water
$\Delta H = 0$ e.g. Benzene-Ether solution	

- Relative lowering of vapour pressure is independent of temperature. Explain? (2 times)

- Ans: Relative lowering of vapour pressure of a solution is equal to the mole fraction of solute.
- $$\frac{\Delta P}{P^\circ} = X_2$$



Vapour pressure ( $p^0$ ) and lowering of vapour pressure ( $\Delta p$ ) changes with temperature. When temperature of solution increases both factors  $\Delta p$  and  $p^0$  increases. Solvent ratio remains same.

### 6. Solubility and solubility curves:

#### 11. Differentiate between continuous and discontinuous solubility curves? (3 times)

**Ans:** Continuous solubility curves: These are smooth curves and do not show any sharp break points. (Continuous increase or decrease in solubility with temperature. e.g. Solubility curves of  $\text{KClO}_3$ ,  $\text{K}_2\text{Cr}_2\text{O}_7$  etc.

**Discontinuous solubility curves:** These are not smooth and show sudden breaks due to sudden changes in solubilities with increase in temperature. e.g. Solubility curves of  $\text{Na}_2\text{SO}_4$ ,  $\text{CaCl}_2$  etc.

#### 12. Fractional crystallization technique is used to purify the chemical products. Justify? (2 times)

**Ans:** The separation of solid substance from a solution one by one on cooling is called fractional crystallization. Solubilities depend upon temperature. e.g. Solubility of  $\text{KNO}_3$  rapidly changes with temperature but solubility of  $\text{KCl}$  and  $\text{KBr}$  change gradually. Thus one substance may precipitate earlier by cooling, leaving behind others.

#### 13. Define solubility curve. Name its two types?

**Ans:** Solubility curve: A graphical representation between temperature and solubility of a substance is called solubility curve. These are of two types:

(i) Continuous solubility curve (ii) Discontinuous solubility curve

#### 14. Why the solubility of glucose increases on increasing temperature?

**Ans:** When glucose is dissolved in water it shows endothermic heat of solution. Solute molecules separate from each other to dissolve in solvent. This process requires energy. So when temperature is increased solubility also increases.

#### 15. Can sugar not be dissolved in benzene. Give reason?

**Ans:** Can-sugar is a polar covalent substance and it is soluble in polar solvent like water. Sugar dissolves in water due to the formation of H-bonding between solute and solvent, but not in benzene. As we simply say that solubility is based upon principle "Like dissolves like".

#### 16. Define solubility giving one example? (2 times)

**Ans:** The solubility is defined as the concentration of the solute in the solution when it is in equilibrium with the solid substance at a particular temperature. Solubility is expressed in terms of number of grams of solute in 100g of solvent. At a particular temperature, saturated solution of  $\text{NaCl}$  in water at  $0^\circ\text{C}$  contains 37.5g of  $\text{NaCl}$  in 100g of water.

### 7. Colligative Properties:

#### 17. Define colligative properties. Name important colligative properties? (8 times)

**Ans:** Colligative properties: The properties of a solution which are based upon the number of solute particles and independent of the nature of solute are called colligative properties. Name of some colligative properties are:

- (i) Lowering of vapour pressure
- (ii) Elevation in boiling point
- (iii) Depression in freezing point
- (iv) Osmotic pressure

#### 18. Give the conditions of colligative properties?

**Ans:** There are following conditions for colligative properties:

(iii) Solute must be non-electrolyte.

#### 19. Why some properties are called colligative properties? (2 times)

**Ans:** Colligative properties are called so because these depend upon the number of solute particles in definite amount of solvent and independent of the nature of solute. For example lowering of vapour pressure of water, caused by the addition of 6 g of urea, 18 g of glucose and 34.2 g of sucrose is same although the solute particles are of different nature but their numbers are same.

#### 20. Colligative properties are obeyed when the solute is non-electrolyte and also when the solutions are dilute. Explain?

**Ans:** Colligative properties depend upon number of solute particles in solution of non-electrolyte. Solute does not ionize in solution and its number of particles remains same. e.g. 1 mole of urea produces  $6.02 \times 10^{23}$  particles in solution. When solution is dilute, less number of solute particles far away from each other and they behave independently and colligative properties are obeyed.

#### 21. Describe that $\text{NaCl}$ and $\text{KNO}_3$ are used to lower the melting point of ice?

**Ans:**  $\text{NaCl}$  and  $\text{KNO}_3$  are used as antifreeze. They depress the freezing point of ice and it melts. This method is used to melt the ice from the roads in winter.

#### 22. Why is the vapour pressure of a solution less than pure solvent?

**Ans:** In pure solvent all the surface of solvent is covered by solvent molecules. But when a solute is added to it from a solution, some surface is occupied by solute particles. Hence escaping tendency of solvent is decreased and thus vapour pressure of solution is also lowered.

#### 23. State ebullioscopic constant and cryoscopic constant? (1 time)

**Ans:** **Ebullioscopic constant:** It is elevation in boiling point which is produced, when 1 mole of solute is dissolved in 1 kg of solvent. It is denoted by  $K_b$ . Its value is  $0.52^\circ\text{C}$  for water.

**Cryoscopic constant:** It is depression in freezing point which is produced when 1 mole of solute is dissolved in 1 kg of solvent. It is denoted by  $K_f$ . Its value is  $1.86^\circ\text{C}$  for water.

#### 24. In summer the antifreeze solutions protect the radiators from boiling over. How?

**Ans:** The boiling point of liquid is increased by the addition of solute such as ethylene glycol in water. Therefore in summer the antifreeze liquids protect the water in car radiators from over boiling. Similarly the addition of a solute in water depresses its freezing point. During winter it protects a car by preventing the liquid from freezing in radiators.

#### 25. Boiling points of liquid are increased when solute is added to them. Give reason? (5 times)

**Ans:** Because the vapour pressure of solvent decreases due to the presence of solute. Therefore, we have to supply high temperature in order to equalize the vapour pressure of solvent to external pressure to boil it. This results in the increase of boiling point of solvent in the presence of solute.

#### 26. What are the names of four major parts of apparatus used in Landsberger's method for elevation of boiling point?

**Ans:** (1). An inner tube with a hole in its side. This tube is graduated.  
(2). A boiling flask which sends the solvent vapours into the graduated tube through a rosehead.  
(3). An outer tube, which receives hot solvent vapours coming from the side hole of the inner tube.  
(4). A thermometer which can read up to  $0.01^\circ\text{K}$ .

#### 27. Beckmann's thermometer is used to note the depression in freezing point. Explain with reason? (2 times)

**Ans:** There is a very small difference between freezing point of pure solvent and its dilute solution. Ordinary thermometer can read up to  $0.5^\circ\text{K}$ . Hence these cannot differentiate



between freezing point of pure solvent and solution. Beckmann's thermometer can read up to 0.01 K. Hence it can exactly measure the freezing point of pure solvent and solution.

28. The freezing points are depressed due to the presence of solutes. Give reasons? (2 times)

Ans: Freezing point of a substance also related to its vapour pressure. It is a temperature at which there is equilibrium between a solid and a liquid. At freezing point, vapour pressure of liquid and solid phase is same. When some solute is added to solvent, its vapour pressure decreases and freezing point also decreases.

29. Define molal freezing point constant giving example.

Ans: It is depression of freezing point, when one mole of non-volatile and non-electrolyte solute is dissolved in one kg of solvent. It is denoted by  $K_f$ .

30. Why ethylene glycol is added in radiator of Automobile?  
Ans: The most important application of this phenomenon is the use of antifreeze in the radiator of an automobile. The solute is ethylene glycol, which is not only completely miscible with water but has a very low vapour pressure and non-volatile in character. When mixed with water it lowers the freezing point as well as raised the boiling point.

31. Write two applications of boiling point elevation and freezing-point depression in daily life.

Ans: (1). The most important application of this phenomenon is the use of an antifreeze in the radiator of an automobile. The solute is ethylene glycol, which is not only completely miscible with water but has a very low vapour pressure and non-volatile in character. When mixed with water it lowers the freezing point as well as raised the boiling point.

(2). Another common application is the use of NaCl or  $\text{KNO}_3$  to lower the melting point of ice. One can prepare a freezing mixture for use in an ice cream machine.

### 8. Energetics of Solution:

32. Why hydration energy of  $\text{Na}^+$  ion is less than  $\text{Li}^+$  ion? (1 time)

Ans: Hydration energy depends upon charge/size ratio. Greater the charge to size ratio, greater would be the value of hydration energy. As  $\text{Li}^+$  ion has greater charge to size ratio than  $\text{Na}^+$  ion. Thus, its hydration energy is greater.

33. Define "heat of solution" and "hydration energy"?

Ans: The quantity of heat energy, that is absorbed or released when a substance forms solution, is termed as heat of solution.

When ionic compound is dissolved in water, the first step is the separation of ions from solid and second step is separated ions are surrounded by solvent molecules. A hydrated ion is attracted by the solvent dipoles and energy is released, so second step is exothermic. The energy given out by the second step is known as the hydration energy.

2019

34. How will you prepare 10% w/v glucose solution.

Ans: For this purpose we dissolve 10 grams of glucose in water to make the volume 100 cm<sup>3</sup>.

35. Write two differences between ideal & non-ideal solutions.

Ans: Ideal solutions:

- These obey Raoult's law.
- These have zero enthalpy of solution.

Non-ideal solutions:

- They do not obey Raoult's law.
- They have negative or positive enthalpy of solution.

Mention two applications of depression of freezing point.

- NaCl or  $\text{KNO}_3$  are used to lower the m.p of ice. NaCl is used in de-icing of roads.
- NaCl is used to prepare freezing mixture for ice creams.

37. Hydration energy of  $\text{Li}^+$  is greater than  $\text{Cs}^+$ .  
Ans:  $\text{Li}^+$  has greater charge/size ratio as compared to  $\text{Cs}^+$ . So therefore it has greater hydration energy.

## CHAPTER-9 (SOLUTIONS) LONG QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

Describe four points of differences between ideal and non-ideal solution?  
What are ideal solutions? Explain the fractional distillation of ideal mixture of two liquids?

How ideal mixture of two liquids can be separated by fractional distillation?  
Explain the fractional distillation. Justify the two curves when composition is plotted against boiling point of the solution?

Explain positive and negative deviation of non-ideal solution. (2 times)  
What is the difference between ideal and non-ideal solution. Give two points for each. Also define Raoult's Law in two ways mathematically? (2 times)

Explain Raoult's law when both components in the solution (solute and solvent) are volatile? (2 times)

Define Raoult's Law and derive its different mathematical forms when solute is Non Volatile and Solvent is Volatile?

State and explain Raoult's law in three forms. (11 times)

Discuss Raoult's law for the solution in which both components are volatile. (1 time)

Explain negative deviation of solution from Raoult's law?

What are azeotropic mixtures? What types of deviation are shown by them. Explain with the help of graphs?

What is azeotropic mixture? Explain them with the help of graphs?  
Define the colligative properties and explain the lowering of vapour pressure is a colligative property?

How lowering of the vapour pressure as colligative property is used to find out molecular mass of solute. (3 times)

Define colligative properties. Explain elevation of the boiling point?

Define solubility and make clear difference between continuous and discontinuous solubility curves? (3 times)

Give graphical explanation for elevation of boiling point of a solution?

Describe the method to determine the elevation of boiling point of a solution? (1 time)

Explain elevation of boiling point with the help of figure?

Write graphical explanation for elevation of boiling point of a solution.  
Define colligative properties and discuss freezing point depression by Beckman's apparatus.

Describe one method to record the depression of freezing point of a solution (2 times)

Derive molecular mass of compound by it?

How molecular mass of an organic compound is determined by Landsberger method?

How can you measure elevation of boiling point by Landsberger's method? (5 times)



## Solved Exercise Chapter # 9

### ALP SMART SYLLABUS 2020

11. Choose the correct answer for the given ones. (C.W)
- i) 18 g glucose is dissolved in 90 g of water. The relative lowering of vapour pressure is equal to  
(a) 1/5 (b) 5.1 (c) 1/51 (d) 6
- ii) An azeotropic mixture of two liquids boils at a lower temperature than either of them when:  
(a) it is saturated (b) it shows positive deviation from Raoult's law  
(c) it shows negative deviation from Raoult's law (d) it is metastable
- (vi) In azeotropic mixture showing positive deviation from Raoult's law, the volume of the mixture is  
(a) slightly more than the total volume of the components  
(b) slightly less than the total volume of the components  
(c) equal to the total volume of the components  
(d) none of these
- (vii) Which of the following solutions has the highest boiling point?  
(a) 5.85 % solution of sodium chloride (b) 18.0 % solution of glucose  
(c) 6.0 % solution of urea (d) All have the same boiling point
- (viii) Two solutions of NaCl and KCl are prepared separately by dissolving same amount of the solute in water. Which of the following statements is true for these solutions?  
(a) KCl solution will have higher boiling point than NaCl solution  
(b) Both the solutions have different boiling points  
(c) KCl and NaCl solutions possess same vapour pressure  
(d) KCl solution possesses lower freezing point than NaCl solution
- (ix) The molal boiling point constant is the ratio of the elevation in boiling point to  
(a) molarity (b) molality  
(c) mole fraction of solvent (d) mole fraction of solute
- (x) Colligative properties are the properties of  
(a) dilute solutions which behave as nearly ideal solutions  
(b) concentrated solutions which behave as nearly non-ideal solutions  
(c) both (i) and (ii) (d) neither (i) nor (ii)
- Q 2. Fill in the blanks with suitable words (C.W)
- (ii) 100g of a 10% aqueous solution of NaOH contains 10g of NaOH in 90 g of water.
- (iii) When an azeotropic mixture is distilled, its composition remains constant.
- (iv) The molal freezing point constant is also known as cryoscopic constant.
- (v) The boiling point of an azeotropic solution of two liquids is lower than either of them because the solution shows positive deviation from Raoult's law.
- (vi) Among equimolar aqueous solutions of NaCl, BaCl<sub>2</sub> and FeCl<sub>3</sub>, the maximum depression in freezing point is shown by FeCl<sub>3</sub> solution.
- (vii) A solution of ethanol in water shows positive deviations and gives azeotropic solution with lower boiling point than other components.
- (viii) Colligative properties are used to calculate molar mass of a compound.
- (ix) The hydration energy of Br<sup>-</sup> ion is less than that of F<sup>-</sup> ion.
- (x) The aqueous solution of NH<sub>4</sub>Cl is acidic while that of Na<sub>2</sub>SO<sub>4</sub> is neutral.
- Q 3. Indicate True or False from the given statements (C.W)
- (i) At a definite temperature the amount of a solute in a given saturated solution is fixed. (True)
- (ii) Polar solvents readily dissolve non-polar covalent compounds. (False)
- (iii) The solubility of a substance decreases with increase in temperature, if the heat of a solution is negative. (True)
- (iv) The rate of evaporation of a liquid is inversely proportional to the intermolecular forces of attraction. (True)

- (v) The molecular mass of an electrolyte determined by lowering of vapour pressure is less than the theoretical molecular mass.
- (vi) Boiling point elevation is directly proportional to the molality of the solution and inversely proportional to boiling point of solvent (True)
- (vii) All solutions containing 1g of non-volatile non-electrolyte solutes in some solvent will have the same freezing point. (True)
- (viii) The freezing point of a 0.05 molal solution of a non-volatile non-electrolyte in water is -0.93°C. (False)
- Q5.(b) One has one molal solution of NaCl and one molal solution of glucose. (C.W)
- (i) Which solution has greater number of particles of solute? (False)
- (ii) How do we convert these concentrations into weight by weight percentage?
- (iii) Since both solutions are one molal so the number of molecule of solute should be same. But actually NaCl is an electrolyte. Each formula unit of NaCl gives two ions and two particles. If NaCl is 100% dissociated in the solution then one molal solution of NaCl in 1kg of water will give twice the Avogadro's number of particles of solutes. Glucose is a non-electrolyte. Its one molal solution in 1kg of water will give Avogadro's number of particles.
- (ii) One mole of NaCl (58.5g) will have a smaller volume than one mole of glucose (180g). Anyhow the amount of solvent in both solutions is 1kg so the amount of the solvent in both solutions is equal.
- (iii) Number of moles of NaCl = 1, Mass of NaCl dissolved = 58.5g  
Mass of solvent water = 1000g  
Total mass of solution = 1000+58.5 = 1058.5g  
Percentage of NaCl = ?  
Formula applied:
- $$\text{Percentage of NaCl} = \frac{\text{Mass of NaCl}}{\text{Total mass of solution}} \times 100$$
- Putting values
- $$\text{Percentage of NaCl} = \frac{58.5}{1058.5} \times 100 = 5.52\%$$
- Now calculate % age of glucose
- Data: Mass of glucose dissolved = 180g, Mass of solvent water = 1000g  
Percentage of glucose = ?  
Total mass of solution = 1000+180 = 1180g  
Percentage of glucose =  $\frac{\text{Mass of glucose dissolved}}{\text{Total mass of solution}} \times 100$
- $$= \frac{180}{1180} \times 100 = 15.2\%$$
- Q8.(a) Explain fractional distillation. Justify the two curves when composition is plotted against boiling point of solutions. (H.W)
- Ans: Consult Text Book Page No. 263-265
- (b) The solutions showing positive and negative deviations cannot be fractionally distilled at their specific compositions. Explain it.
- Ans: Consult Text Book Page No. 264-265
- Q9.(a) What are azeotropic mixtures? Explain them with the help of graphs? (H.W)
- Ans: Consult Text Book Page No. 264
- (b) Explain the effect of temperature on phenol-water system.
- Ans: Consult Text Book Page No. 259
- Q10.(a) What are colligative properties? Why are they called so? (H.W)
- Ans: Consult Text Book Page No. 267-268
- (b) What is the physical significance of  $K_b$  and  $K_f$  values of solvents?
- Ans: Consult Text Book Page No. 268



Q11. How do you explain that the lowering of vapour pressure is a colligative property? How do we measure the molar mass of a non volatile, non-electrolyte solute in a volatile solvent? (H.W)

Ans: Consult Text Book Page No. 268-269 (C.W)

Q12. How do you justify that

- boiling points of the solvents increase due to the presence of solutes.
- freezing points are depressed due to the presence of solutes.
- the boiling point of one molal urea solution is 100.52 °C but the boiling point of two molal urea solution is less than 101.04°C.
- Beckmann's thermometer is used to note the depression in freezing point.
- in summer the antifreeze solutions protect the liquid of the radiator from boiling over.

Ans(a) NaCl and KNO<sub>3</sub> are used to lower the melting point of ice. The surface of the solution has molecules of solute as well. They do not allow the solvent to leave the surface as rapidly as in pure solvent. To boil the solutions, we have to increase the temperature of solutions in comparison to pure solvents. The B.P of solution are higher than pure solvents.

(b) The lowering of vapour pressure compels the solution to freeze at those temperatures, which are below the freezing point of pure solvent. The reason is that the vapour pressure meets the solid phase of pure solvent at lower temperature than the pure solvent.

(c) 2 molal urea solution is more concentrated than 1 molal. Greater number of molecules of solute cannot make their molecules so apart from each other that there are twice number of particles (molecules). Some molecule make dimers due to less space available in water. Thus in 2 molal solution there are no double the number of Avogadro's number of particles. So colligative property is not doubled. Beckman thermometer can measure up to 1/20<sup>th</sup> of the degree. The elevation of boiling points and the depression of freezing points for dilute solutions are very small quantities. Hence one, can measure these very small changes of temperatures.

(e) Water boils at 100°C. It is used in the radiators to decrease the temperature of the working engine. If we add some suitable solutes which increase the boiling point of water, above 100°C, then easy boiling over of water is avoided. Actually such solutes also decrease the F.P of solutions as well.

(f) NaCl and KNO<sub>3</sub> are electrolyte and are sufficiently soluble in water. The double the number of particles after dissociation in water. In this way they can manage to decrease the freezing point of water to a greater extent as compared to a non-electrolyte.

Q13. What is Raoult's law. Give its three statements. How this law can help us to understand the Ideality of a solution. (H.W)

Ans: Consult Text Book Page No. 260-261

Q14. Give graphical explanation for elevation of boiling point of a solution. Describe one method to determine the boiling point elevation of a solution. (H.W)

Ans: Consult Text Book Page No. 270-272

Q15. Freezing points of solutions are depressed when non-volatile solutes are present in volatile solvents. Justify it. Plot a graph to elaborate your answer. Also, give one method to record the depression of freezing point of a solution (H.W)

Ans: Consult Text Book Page No. 272-274

Q21. 4.675g of a compound with empirical formula C<sub>6</sub>H<sub>5</sub>O were dissolved in 212.5 g of pure benzene. The freezing point of solution, was found 1.02°C less than that of pure benzene. The molal freezing point constant of benzene is 5.1°C. Calculate (i) the relative molar mass and (ii) the molecular formula of the compound. (C.W)

Ans: Data:

Mass of solute dissolved (W <sub>2</sub> )	=	4.675g
Mass of solvent benzene (W <sub>1</sub> )	=	212.5g
Depression of F.P. of solution (ΔT <sub>f</sub> )	=	1.02°C
Molal F.P constant of solvent (K <sub>f</sub> )	=	5.1°C

Relative molar mass of solute (M<sub>1</sub>)

The molecular formula of the compound

Formula applied:

$$M_1 = \frac{K_f \cdot 1000W_2}{\Delta T_f \cdot W_1}$$

Putting the values of various factors.

$$M_1 = \frac{5.1^\circ\text{C} \times 1000 \times 4.675\text{g}}{1.02^\circ\text{C} \times 212.5\text{g}} = \frac{5.1 \times 1000 \times 4.675}{1.02 \times 212.5} = \frac{23842.5}{216.75} = 110\text{g mol}^{-1}$$

Empirical formula

Empirical formula mass

Formula applied:

Molal mass

Putting values

$$\begin{aligned} &= \text{C}_6\text{H}_5\text{O} \\ &= 36 + 5 + 16 = 55 \\ &= n(\text{E.F. mass}) \end{aligned}$$

$$\begin{aligned} 110 &= n(55) \\ n &= \frac{110}{55} = 2 \\ &= 2(\text{C}_6\text{H}_5\text{O}) \end{aligned}$$

Molecular formula of the compound =

Hence the molecular formula of the compound = C<sub>12</sub>H<sub>10</sub>O<sub>2</sub>

Q22. The boiling point of a solution containing 0.2 g of a substance A in 20.0 g of ether (molar mass = 74) is 0.17 K higher than that of pure ether. Calculate the molar mass of A. Molal boiling point constant of ether is 2.16 K. (C.W)

Ans:

Mass of the solute A (W <sub>2</sub> )	=	0.2g
Mass of the solvent ether (W <sub>1</sub> )	=	20.0g
Elevation of B.P of solution (ΔT <sub>b</sub> )	=	0.17 K
Molal B.P constant of ether (K <sub>b</sub> )	=	2.16K
Molar mass of A (M <sub>2</sub> )	=	?

Formula applied:

$$M_2 = \frac{K_b \times 1000W_2}{\Delta T_b \cdot W_1}$$

Putting the values of various factors.

$$M_2 = \frac{2.16\text{K} \times 1000 \times 0.2\text{g}}{0.17\text{K} \times 20\text{g}} = \frac{132}{3.4} = 127.0\text{g mol}^{-1}$$

Q23. 3 g of a non-volatile, non-electrolyte solute 'X' are dissolved in 50 g of ether (molar mass = 74) at 293 K. The vapour pressure of ether falls from 442 torr to 426 torr under these conditions. Calculate the molar mass of solute 'X'. (C.W)

Data: Mass of the solute 'X' dissolved (W <sub>2</sub> )	=	0.2g
Mass of the solvent ether (W <sub>1</sub> )	=	20.0g
Molar mass of ether (M <sub>1</sub> )	=	74.00 g mol <sup>-1</sup>
Vapour pressure of pure ether (P <sup>o</sup> )	=	442 torr
Vapour pressure of solution (P')	=	426 torr
Lowering of vapour pressure (ΔP)	=	442 - 426 = 16 torr
Molal B.P constant of ether (K <sub>b</sub> )	=	2.16K
Molar mass of X (M <sub>2</sub> )	=	?

Formula applied:

$$\frac{\Delta P}{P^o} = \frac{W_2}{M_2} \times \frac{M_1}{W_1}$$

$$M_2 = \frac{W_2 \times M_1 \times P^o}{\Delta P \times W_1}$$

Putting the values of various factors.

$$M_2 = \frac{3.0\text{g} \times 74\text{g mol}^{-1} \times 442\text{torr}}{16\text{torr} \times 50.0\text{g}} = \frac{13 \times 74 \times 442}{16 \times 50} = \frac{98124}{800} = 122.6\text{g mol}^{-1}$$



**CHAPTER -10 (ELECTROCHEMISTRY)**  
**OBJECTIVES (MCQ'S)**  
**(ACCORDING TO ALP SMART SYLLABUS-2020)**

- In  $H_2SO_4$  the oxidation state of oxygen is:  
 (a) +1 (b) -1 (c) +2 (d) -2
- In  $Na_2O_2$  the oxidation state of oxygen is:  
 (a) -2 (b) +2 (c) -1 (d) +1
- The oxidation number of nitrogen in  $HNO_3$  is:  
 (a) +3 (b) -3 (c) -5 (d) +5
- The oxidation number of chromium in  $K_2Cr_2O_7$  is:  
 (a) +14 (b) +12 (c) +6 (d) +13 (4 times)
- Oxidation number of phosphorous in the compound  $(HPO_3)_2$  is:  
 (a) +3 (b) +4 (c) +5 (d) +6
- Oxidation number of "Cr" in  $K_2CrO_4$  is:  
 (a) +2 (b) +4 (c) +6 (d) +8 (2 times)
- Oxidation number of oxygen in  $OF_2$  is:  
 (a) Zero (b)  $-\frac{1}{2}$  (c) +2 (d) -1 (3 times)
- The oxidation number of Cl in  $HClO_4$  is:  
 (a) +2 (b) +3 (c) +5 (d) +7
- Oxidation number of Cl in  $Ca(ClO_3)_2$  is:  
 (a) -1 (b) +5 (c) +3 (d) +1
- Oxidation state of Mn in  $MnO_4^{2-}$  is:  
 (a) +4 (b) +6 (c) +5 (d) -6 (1 time)
- The process during which one metal is deposited on the surface of an other metal by using electrical current is called:  
 (a) Electrolysis (b) Electroplating (c) Electrolytic refining (d) Electrolytic Purification
- Which process is used for the extraction of Al?  
 (a) Caster-Kelliner process (b) Thermite process (c) Hall-Beroult (d) Combustion process
- An oxidizing agent:  
 (a) Loses electrons (b) is oxidized  
 (c) Neither gain nor loses electrons (d) Gains electrons
- The electrolysis of aqueous solution of  $NaNO_3$  at cathode liberates  
 (a) Sodium metal (b)  $H_2$  gas (c)  $O_2$  gas (d)  $NO_2$  gas
- Galvanic cells which cannot be recharged are called:  
 (a) Diffused cell (b) Secondary cells (c) Tertiary cells (d) Primary cells
- Standard Hydrogen Electrode (SHE) is made of:  
 (a) Ag foil (b) Au foil (c) Cu foil (d) Pt foil (2 times)
- Electrode potential of SHE arbitrarily taken in volts is:  
 (a) 0.00 (b) 1.00 (c) 0.01 (d) 0.50
- Reduction always take place at:  
 (a) Anode (b) Cathode (c) Both electrodes (d) Salt bridge
- The highest reduction potential in electrochemical series is of  $F_2$  and its value is:  
 (a) -3.87V (b) -2.87V (c) +2.87V (d) +3.87V
- A Single lead cell provides volts:  
 (a) 2 (b) 4 (c) 6 (d) 8
- The best reducing agent is:  
 (a)  $F^{-1}$  (b)  $Cl^{-1}$  (c)  $Br^{-1}$  (d)  $I^{-1}$
- The oxidation state of oxygen in  $KO_2$  is:  
 (A) -1 (B) -2 (C) -1/2 (D) +2
- The number of moles of hydrogen atoms in 92 g alcohol ( $C_2H_5OH$ ) are:  
 (A) 5 moles (B) 6 moles (C) 10 moles (D) 12 moles
- If salt bridge is not used between two half cells then the voltage:  
 (A) decreases rapidly (B) decreases slowly (C) does not change (D) drops to zero
- Oxidation number of chromium in  $Cr_2O_3$  is:  
 (A) +1 (B) +2 (C) +3 (D) +4

The oxidation state of Mn in  $KMnO_4$  is:  
 (b) +6 (c) +2 (d) +5

Answers													
1	2	3	4	5	6	7	8	9	10	11	12	13	14
D	C	D	C	C	C	C	D	B	B	B	C	D	B
15	16	17	18	19	20	21	22	23	24	25	26		
D	D	A	B	C	A	A	C	D	D	C	A		

**CHAPTER-10 (ELECTROCHEMISTRY)**  
**SHORT QUESTIONS**  
**(ACCORDING TO ALP SMART SYLLABUS-2020)**

- Define electrochemistry?  
 It is the branch of chemistry which deals with the conversion of electrical energy into chemical energy in electrolytic cells as well as the conversion of chemical energy into electrical energy in galvanic or voltaic cell. (4 times)
- Oxidation State and Balancing of Redox Equation:  
 Calculate oxidation number of N in  $HNO_3$  and  $NO_2$ ?  
 Ans:  $HNO_3$ :  
 O.N of H = +1  
 O.N of O = -2  
 O.N of N = ?  
 (+1) + (Oxidation No. of N) + 3(-2) = 0  
 (Oxidation No. of N) - 5 = 0  
 (Oxidation No. of N) = 5  
 $NO_2$ :  
 O.N of H = +1  
 O.N of O = -2  
 O.N of N = ?  
 (Oxidation No. of N) + 2(Oxidation No. of O) = 0  
 (Oxidation No. of N) + 2(-2) = 0  
 (Oxidation No. of N) - 4 = 0  
 (Oxidation No. of N) = +4
- Define oxidation state with two examples? (3 times)  
 Ans: Oxidation state: It is the apparent charge on an atom of an element in a molecule or an ion. It may be positive or negative. e.g. Oxidation No. of Mn is +7 in  $KMnO_4$ :  
 O.N of K = +1  
 O.N of O = -2  
 O.N of Mn = ?  
 (+1) + (Oxidation No. of Mn) + 4(-2) = 0  
 (+1) + (Oxidation No. of Mn) - 8 = 0  
 (Oxidation No. of Mn) - 7 = 0  
 (Oxidation No. of Mn) = +7  
 $HNO_3$ :  
 (Oxidation No. of N) + 3(Oxidation No. of O) = 0  
 (+1) + (Oxidation No. of N) + 3(-2) = 0  
 (Oxidation No. of N) - 5 = 0  
 (Oxidation No. of N) = +5

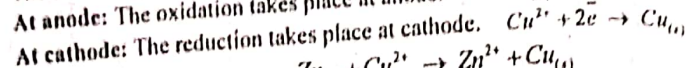
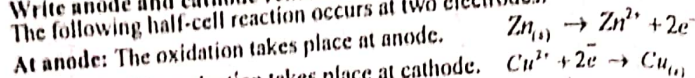


$$(+1) + (\text{Oxidation No. of N}) - 6 = 0$$

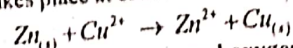
$$(\text{Oxidation No. of N}) - 5 = 0$$

$$(\text{Oxidation No. of N}) = +5$$

4. Write anode and cathode reactions of a galvanic cell?  
 Ans: The following half-cell reaction occurs at two electrodes.



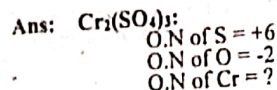
Overall reaction:



5. Give rules for oxidation number of hydrogen and oxygen?  
 Ans: (i) Oxidation number hydrogen is +1, except in case of metal hydride. Where it is -1, e.g. of  $\text{H}^+\text{Cl}^-$ ,  $\text{Na}^+\text{H}^-$ .

(ii) Oxidation number of oxygen is -2 except in case of per oxides and super oxides where it is -1 and  $-\frac{1}{2}$  respectively.

- (iii) Oxidation No. of oxygen is +2 in  $\text{OF}_2$ .  
 6. Calculate the oxidation number of chromium in: (a)  $\text{Cr}_2(\text{SO}_4)_3$  (b)  $\text{K}_2\text{Cr}_2\text{O}_7$  (5 times)



$$\text{Cr}_2(\text{SO}_4)_3$$

$$2(\text{Oxidation No. of Cr}) + 3[(\text{Oxidation No. of S}) + 4(\text{Oxidation No. of O})] = 0$$

$$2(\text{Oxidation No. of Cr}) + 3[(+6) + 4(-2)] = 0$$

$$2(\text{Oxidation No. of Cr}) + 3[+6 - 8] = 0$$

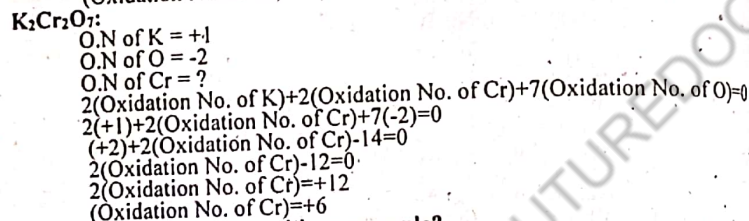
$$2(\text{Oxidation No. of Cr}) + 3[-2] = 0$$

$$2(\text{Oxidation No. of Cr}) + 3[-2] = 0$$

$$2(\text{Oxidation No. of Cr}) - 6 = 0$$

$$2(\text{Oxidation No. of Cr}) = +6$$

$$(\text{Oxidation No. of Cr}) = +3$$

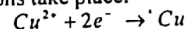


7. Define oxidation reaction with an example?

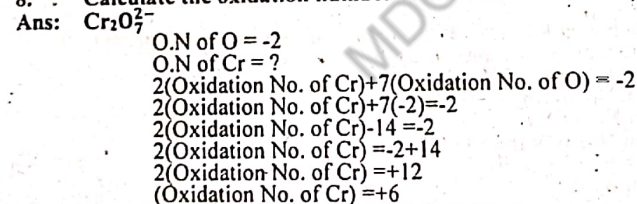
Ans: Oxidation reaction: Those reaction in which increase in oxidation state or loss of electrons take place.



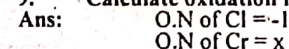
Reduction reaction: Those reactions in which decrease in oxidation number or gain of electrons take place.



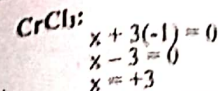
8. Calculate the oxidation number of chromium in:  $\text{Cr}_2\text{O}_7^{2-}$  (2 times)



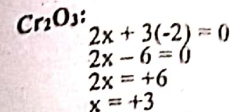
9. Calculate oxidation number of Cr in  $\text{CrCl}_3$ ?



(2 times 2018)

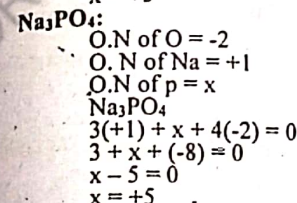
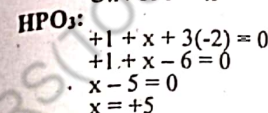


10. Calculate the oxidation number of chromium in  $\text{Cr}_2\text{O}_3$ ?  
 Ans: O.N of oxygen = -2  
 O.N of Cr = x (1 time) (2017 = 1 time)

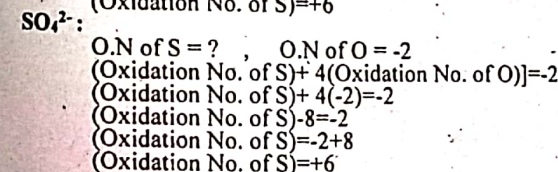
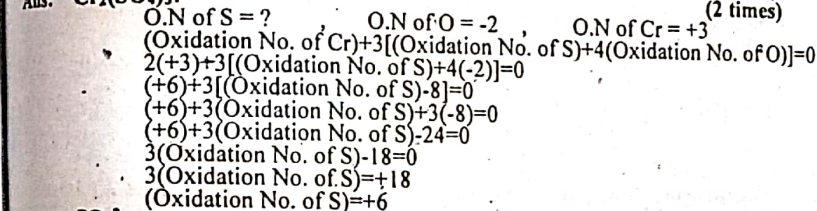


11. Give four rules for assigning of oxidation number?  
 Ans: (i) The oxidation number of free elements is zero. For example H, Mg, Na.  
 (ii) The oxidation number of an ion consisting of a single element is the same as charge on the ion.  
 (iii) Oxidation number of hydrogen in all its compounds is +1 except metal hydride where it is -1.  
 (iv) In neutral molecules, the algebraic sum of oxidation number of all the elements is zero.

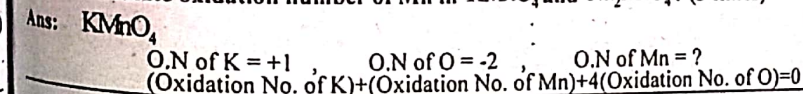
12. Calculate the oxidation number of phosphorous in (a)  $\text{HPO}_3$  (b)  $\text{Na}_3\text{PO}_4$ ?  
 Ans:  $\text{HPO}_3$ :  
 O.N of oxygen = -2  
 O.N of hydrogen = +1  
 O.N of P = x



13. Calculate oxidation number of 'S' in: (i)  $\text{Cr}_2(\text{SO}_4)_3$  (ii)  $\text{SO}_4^{2-}$  (2 times)



14. Calculate oxidation number of Mn in  $\text{KMnO}_4$  and  $\text{Na}_2\text{MnO}_4$ ? (3 times)





$$(+2) + (\text{Oxidation No. of S}) - 8 = 0$$

$$(\text{Oxidation No. of S}) - 6 = 0$$

$$(\text{Oxidation No. of S}) = +6$$

Calculate the oxidation state of underlined element. (a)  $\text{H}_3\text{P}\underline{\text{O}}_3$  (b)  $\text{Ca}(\underline{\text{ClO}}_3)_2$

20. Ans:  $\text{H}_3\text{P}\underline{\text{O}}_3$ :  
 $3(\text{Oxidation No. of H}) + (\text{Oxidation No. of P}) + 3(\text{Oxidation No. of O}) = 0$   
 $3(+1) + (\text{Oxidation No. of P}) + 3(-2) = 0$   
 $(+3) + (\text{Oxidation No. of P}) - 6 = 0$   
 $\text{Oxidation No. of P} - 3 = 0$   
 $\text{Oxidation No. of P} = +3$

$\text{Ca}(\underline{\text{ClO}}_3)_2$ :  
 $(\text{Oxidation No. of Ca}) + 2[(\text{Oxidation No. of Cl}) + 3(\text{Oxidation No. of O})] = 0$   
 $(+2) + 2[(\text{Oxidation No. of Cl}) + 3(-2)] = 0$   
 $(+2) + 2(\text{Oxidation No. of Cl}) + 6(-2) = 0$   
 $(+2) + 2(\text{Oxidation No. of Cl}) - 12 = 0$   
 $2(\text{Oxidation No. of Cl}) - 10 = 0$   
 $2(\text{Oxidation No. of Cl}) = +10$   
 $(\text{Oxidation No. of Cl}) = +5$

21. Ans: Calculate Oxidation Number of Cr in : (i)  $\text{K}_2\text{CrO}_4$  (ii)  $\text{Cr}_2\text{O}_3$  (2 times 2018)

$\text{K}_2\text{CrO}_4$ :  
 $2(\text{Oxidation No. of K}) + (\text{Oxidation No. of Cr}) + 4(\text{Oxidation No. of O}) = 0$   
 $2(+1) + (\text{Oxidation No. of Cr}) + 4(-2) = 0$   
 $(+2) + (\text{Oxidation No. of Cr}) - 8 = 0$   
 $(\text{Oxidation No. of Cr}) - 6 = 0$   
 $(\text{Oxidation No. of Cr}) = +6$

$\text{Cr}_2\text{O}_3$ :  
 $2(\text{Oxidation No. of Cr}) + 3(\text{Oxidation No. of O}) = 0$   
 $2(\text{Oxidation No. of Cr}) + 3(-2) = 0$   
 $2(\text{Oxidation No. of Cr}) - 6 = 0$   
 $2(\text{Oxidation No. of Cr}) = +6$   
 $(\text{Oxidation No. of Cr}) = +3$

## Electrolytic Conduction:

22. Ans: Differentiate between electrolytic and voltaic cells? (5 times)

	Electrolytic cell	Voltaic cell
(i)	A cell in which electrical energy is converted into chemical energy.	(i) A cell in which chemical energy is converted into electrical energy.
(ii)	Non-spontaneous redox reaction takes place here.	(ii) Spontaneous redox reaction takes place here.
(iii)	Example: Down's cell	(iii) Example: Daniel cell

23. Ans: Differentiate between a cell and battery?

**Cell:** The arrangement in which electrical energy is converted into chemical energy or chemical energy is converted into electrical energy is called cell. It consists of two electrodes and an electrolyte.

**Battery:** The combination of two or more cells is called battery.

24. Ans: Differentiate between electrolysis and electrolytic conduction? (1 time)

**Electrolysis:** The process in which electricity is used to carry out a non-spontaneous redox reaction is called electrolysis.

**Electrolytic conduction:** The conduction of electricity carried out by ions present in fused or aqueous solution of an electrolyte is called electrolytic conduction.

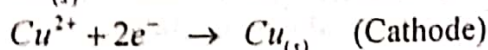
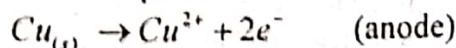
25. Ans: Define electrolytic cell. Give example?

**Electrolytic cell:** A cell in which electric current is used to carry out a non-spontaneous reaction is called electrolytic cell.

Example: (i) Down's cell (ii) Nelson cell.

26. Ans: Impure copper can be purified by electrolytic process. Explain it? (6 times)

During purification of Cu, impure Cu is made anode and thin sheet of pure Cu is made cathode and  $\text{CuSO}_4$  solution is taken as electrolyte. Cu atoms at anode lose electrons and go into the solution as  $\text{Cu}^{2+}$  ion. They migrate to cathode where they gain electrons and are deposited at cathode. In this way Cu atoms migrate from anode to cathode and impurities are left in electrolyte.





27. Give the chemistry of electrolysis of aqueous solution of sodium chloride?  
 Ans: Caustic soda is obtained on industrial scale by the electrolysis of concentrated aqueous solution of sodium chloride using titanium anode and mercury or steel cathode. This electrolysis is carried out in Nelson cell and Castner Kellner cell or Hg-cell.  

$$\text{NaCl}(s) \rightleftharpoons \text{Na}^+(aq) + \text{Cl}^-(aq)$$
  
 At anode:  $2\text{Cl}^-(g) \rightarrow \text{Cl}_2(g) + 2e^-$  (oxidation)  
 At cathode:  $2\text{H}_2\text{O}(l) + 2e^- \rightarrow \text{H}_2(g) + 2\text{OH}^-(aq)$  (reduction)  
 Overall reaction:  

$$2\text{Na}^+(aq) + 2\text{Cl}^-(aq) + 2\text{H}_2\text{O}(l) \rightleftharpoons \text{Cl}_2(g) + \text{H}_2(g) + 2\text{Na}^+(aq) + 2\text{OH}^-(aq)$$
28. How anodized aluminum is produced and why it can absorb dyes? (3 times)  
 Ans: Anodized aluminum is prepared by making it an anode in an electrolytic cell containing sulphuric acid or chromic acid, which coats a thin layer of oxide on it. The aluminum oxide layer resists attack by corrosive agents. The freshly anodized aluminum is hydrated and can absorb dyes.
29. What is the difference between metallic conduction and electrolytic conduction?  
 Ans: Electrolytic conduction: The conduction of electricity carried out by ions present in fused or in an aqueous solution of an electrolyte is called electrolytic conduction. Metallic conduction: The conduction of electricity through a metal due to free electrons is called metallic conduction.
30. What is the difference between conduction through metals and molten electrolytes?  
 Ans: Metals are conductors of electricity because of the relatively free movement of their electrons throughout the metallic lattice. This electronic conduction is simply called metallic conduction. Electrolytes in the form of solution in the fused state have the ability to conduct electricity. In this case the current is not carried by free electrons through the solution or through the fused electrolyte. Here, the current is carried by ions having positive and negative charges, these ions are produced in the solution or in fused state due to ionization of the electrolyte.
31. How impure copper can be purified by electrolysis? (4 times)  
 Ans: Refining of copper:  
 Refining of copper is carried out in electrolytic cell. Impure copper acts as anode and pure copper plate acts as cathode. Oxidation takes place at anode as:  

$$\text{Cu}_{(s)} \rightarrow \text{Cu}^{2+}_{(aq)} + 2e^-$$
  
 Reduction reaction takes place at the cathode. Copper ions present in the solution are attracted to the cathode. Where they gain electrons from the cathode and become neutral and deposit on the cathode.  

$$\text{Cu}^{2+}_{(aq)} + 2e^- \rightarrow \text{Cu}_{(s)}$$
 (2 times)
32. What is electrolysis? Give example?  
 Ans: The electrochemical reactions that occur at the electrodes during the electrolytic conduction constitute the phenomenon of electrolysis. For example the electrolysis of brine in electrolytic cell is given as:  

$$\text{NaCl}(s) \rightleftharpoons \text{Na}^+(aq) + \text{Cl}^-(aq)$$
  
 At anode:  $2\text{Cl}^-(g) \rightarrow \text{Cl}_2(g) + 2e^-$  (oxidation)  
 At cathode:  $2\text{H}_2\text{O}(l) + 2e^- \rightarrow \text{H}_2(g) + 2\text{OH}^-(aq)$  (reduction)  
 Overall reaction:  

$$2\text{Na}^+(aq) + 2\text{Cl}^-(aq) + 2\text{H}_2\text{O}(l) \rightleftharpoons \text{Cl}_2(g) + \text{H}_2(g) + 2\text{Na}^+(aq) + 2\text{OH}^-(aq)$$
33. A salt bridge maintains the electrical neutrality in the galvanic cell. Explain? (11 times)  
 Ans: Two half cells are electrically connected by a salt bridge. Consider a Zn-Cu cell. During reactions of this cell, Zn half-cell continuously loses electrons. Thus this positive charge is increasing.  

$$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$$
  
 While Cu receiving electrons.  

$$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$$
  
 Collection of positive charge in Zn electrode and negative charge in Cu would stop the reaction. Salt bridge prevents the net accumulation of charges in either beaker. Thus through salt bridge negative ions move toward Zn half-cell. In this way salt bridge maintain the two solutions electrically neutral.
34. Explain the function of salt bridge? (12 times)  
 Ans: Function of salt bridge: A salt bridge performs following functions:  
 (i) It brings electrical contact between two half cells.  
 (ii) It maintains electrical neutrality of two solutions.

- Define oxidizing agent and reducing agent.  
 Oxidizing Agent (Oxidants):- The substance which help the oxidation to occur are called oxidizing agent or oxidants. They oxidize the other substance by taking electrons from them and get reduced itself.  
 Examples: 1. Acidified  $\text{KMnO}_4$  and  $\text{K}_2\text{Cr}_2\text{O}_7$   
 2. Mostly non-metals are oxidants
- Reducing Agents (Reductants):- The substances which helps the reduction to occur are called reducing agents. They reduce the other substances by giving electrons to them and themselves get oxidized.  
 Examples: 1. Almost all metals are good reductants.  
 2.  $\text{CO}$ ,  $\text{SO}_2$ ,  $\text{H}_2$ ,  $\text{KI}$  are good reductants.
- Electrode Potential:  
 Define standard electrode potential? (14 times)  
 Standard electrode potential: The potential setup when an electrode is in contact with one molar solution of its own ions at  $298\text{K}$  is called standard electrode potential. It is denoted by  $E^\circ$
- Write the importance of standard hydrogen electrode?  
 Standard hydrogen electrode (SHE) is used to determine the electrode potential of other electrode. It is used as reference electrode and its value is  $0.0$  volt. From SHE we drive electrochemical series.  
 SHE act as anode when connected with Cu electrode but as cathode when connected with Zn-electrode. Justify? (5 times)  
 When SHE is coupled with Cu electrode, Hydrogen give electrons to Cu and get oxidized.  

$$\text{H}_2 \rightarrow 2\text{H}^+ + 2e^-$$
  
 Because reduction potential of Cu is greater than hydrogen. Since hydrogen is oxidized, therefore it acts as anode. On the other hand when SHE is connected with Zn-electrode, Zn give electrons to hydrogen electrode and get oxidized.  

$$\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$$
  
 Since Zn have greater tendency to donate electrons than hydrogen. Therefore Zn acts as anode.
- What is meant by electromotive force (emf) of cell?  
 Electromotive force (emf): The measurement of energy that causes current to flow through a circuit is called emf. It is also called voltage.
- Give cathodic and anodic reactions of electrolysis of concentrated aqueous solution of sodium chloride? (2 times)  
 Ans: 
$$2\text{NaCl} \xrightarrow{\text{H}_2\text{O}} 2\text{Na}^+ + 2\text{Cl}^-$$
  
 At anode:  

$$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^- \text{ (Oxidation)}$$
  
 At cathode:  

$$2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2\text{OH}^- \text{ (Reduction)}$$
  
 Overall reaction:  

$$2\text{Na}^+ + 2\text{Cl}^- + 2\text{H}_2\text{O} \rightarrow 2\text{Na}^+ + 2\text{OH}^- + \text{H}_{2(g)} + \text{Cl}_{2(g)}$$
41. Write electrode reactions of electrolysis of fused sodium chloride?  
 Ans: 
$$\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$$
  
 At anode: 
$$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-$$
  
 At cathode: 
$$2\text{Na}^+ + 2e^- \rightarrow 2\text{Na}_{(s)}$$
  
 Net reaction: 
$$2\text{Na}^+ + 2\text{Cl}^- \rightarrow 2\text{Na} + \text{Cl}_2$$
42. Write down the equations for electrode processes in the electrolysis of fused  $\text{PbCl}_2$ ?  
 Ans: 
$$\text{PbCl}_2 \rightarrow \text{Pb}^{2+} + 2\text{Cl}^-$$
  
 At cathode: 
$$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}_{(s)}$$
  
 At anode: 
$$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-$$
  
 Overall reaction: 
$$\text{PbCl}_2 \rightarrow \text{Pb}_{(s)} + \text{Cl}_{2(g)}$$

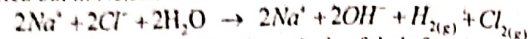


43. Write down the industrial applications of electrolysis. (2 times)

Ans: (i) Na is produced in Down's cell commercially by the electrolysis of fused NaCl by iron cathode and graphite anode.  $\text{Cl}_2$  gas obtained as a by produce.



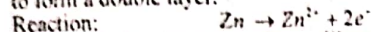
(ii) Caustic soda is produced by electrolysis of an aqueous soln. of NaCl. This process is carried out in Nelson's cell.



(iii) Mg and Ca metals are extracted by electrolysis of their fused chlorides.

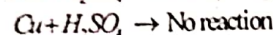
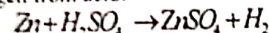
44. Define Zinc electrode potential, Give example? (3 times)

Ans: When a Zn rod is placed in zinc sulphate solution. It bears accumulation of negative charges. This is due to net ionization of some of its atoms. The negative charge of zinc rod will attract an atmosphere of positively charged zinc ions around the rod to form a double layer.



45. Zn can displace hydrogen from dilute acid solution but copper cannot. Justify the statement? (2 times)

Ans: Metals having greater values of reduction potential have less tendency to displace hydrogen from acid. Cu has high value of reduction potential and are not able to displace  $\text{H}_2$ . But Zn has low value of reduction potential and can easily displace hydrogen from acid.



3. Electrochemical Series:

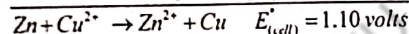
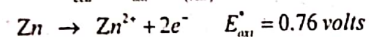
46. Define the electrochemical series?

Ans: The elements are arranged in the order of their standard electrode potentials on the hydrogen scale, the resulting list is known as electrochemical series. This list have been prepared by comparison with standard hydrogen electrode (SHE). In this list, elements above SHE have negative reduction potential while below have positive reduction potential.

47. Give two examples of electrochemical series? (2 times)

Ans: (i) Prediction of the feasibility of a chemical reaction. By the use of oxidation and reduction potential spontaneity of a cell reaction can be determined.

$$\text{Formula: } E_{\text{cell}} = E_{\text{oxi}} + E_{\text{red}}$$



This reaction take place on its own.

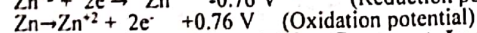
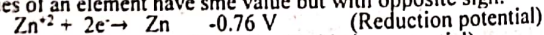
(ii) Relative chemical reactivity of metals can be determined by using electrochemical series, i.e. greater the reduction potential, lesser is the ability to lose electrons. Hence its reactivity is less.

48. Na and K can displace hydrogen from acids but Pt, Pd and Cu cannot. Comment on it? (7 times) (2018)

Ans: Greater the value of standard reduction potential, lesser is the ability to lose electrons. Hence its reactivity is less. Na and K are highly reactive metals. They can easily displace hydrogen from acid whereas Pt, Pd and Cu are least reactive metal. Therefore cannot easily displace hydrogen.

49. Why standard oxidation potential of Zn is 0.76 V and its reduction potential is -0.76V? (2 times)

Ans: The tendency of substance to accept electrons is called its reduction potential and loss of electron is called oxidation potential. Both oxidation and reduction potential values of an element have same value but with opposite sign.



50. Na and K can displace hydrogen from acids but Cu cannot. Justify it? (2 times)

Ans: Greater the value of standard reduction potential of a metal, lesser is its tendency to lose electrons to form metal ions and so weaker is its tendency to displace  $\text{H}_2$ .

from acids. For example, metals like Au, Pt, Ag and Cu which have sufficiently high positive values of reduction potentials, do not liberate hydrogen from acids. While, metals like Na, Mg and K which are close to the top of the series and have very low reduction potentials, liberate hydrogen gas, when they react with acids.

Describe the relative chemical reactivity of metals?

Metals react by forming their positive ions. If the value of standard reduction potential is greater, the tendency to form positive ions is smaller and therefore less reactivity. It means the elements with greater reduction potential are least reactive where as the elements with smaller reduction potential are more reactive.

Zn can displace iron from its solution, how?

Ans: Zn is a stronger reducing agent having standard reduction potential as (0.76) than iron which has standard reduction potential as (0.44). So, Zn can displace iron from its solution.

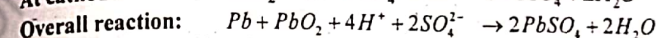
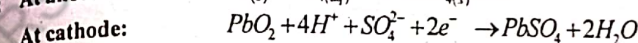
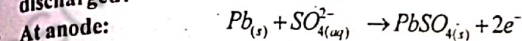
What is Anodized Aluminum?

Ans: When a thin layer of oxide is formed over the surface of aluminium metal it is called anodized aluminium. Anodized aluminium is prepared by making it anode in an electrolyte cell containing sulphuric acid or chromic acid, which coats a thin layer of oxide on it. Aluminium oxide layer resist the attack by corrosive agents.

How relative chemical reactivity of metals is studied with the help of electrochemical series. (2 times)

Ans: The value of the reduction potential of a metal or a non metal tells us the tendency to lose electrons and act a reducing agent. It also gives the information about the tendency of a specie to gain electrons and act as oxidizing agent. Greater the value of standard reduction potential of a given specie, greater is its tendency to accept electrons to undergo reduction and hence to act as oxidizing agent.

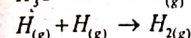
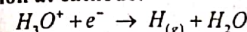
Give the reaction involved, when a lead accumulator (lead storage battery) is discharged? (6 times)



56. How the electrolysis of aqueous solution of sodium nitrate give  $\text{H}_2$  gas at cathode and  $\text{O}_2$  gas at anode respectively? (2 times)

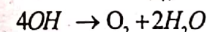
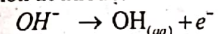
Ans: Consider aqueous solution of  $\text{NaNO}_3$  under electrolysis.

Reaction at cathode:



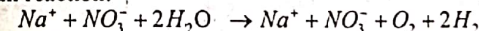
Thus  $\text{H}_2$  gas liberate at cathode.

Reaction at anode:



Hence  $\text{O}_2$  gas discharged at anode.

Overall reaction:



57. The oxidation state of oxygen is +2 in  $\text{OF}_2$ . Justify it.

Ans: The oxidation number of hydrogens in all its compounds except metal hydrides is +1. In metal hydrides it is -1, e.g.  $\text{Na}^+ \text{H}^-$ ,  $\text{Mg}^{2+} \text{H}_2^{(-)2}$ .

$\text{OF}_2$

(Oxidation number of O) + 2 (Oxidation number of F) = 0

(oxidation number of O) + 2 (-1) = 0

(oxidation number of O) - 2 = 0

(oxidation number of O) = +2

58. Write down the difference between ionization and electrolysis.

Ans:	Ionization	Electrolysis
1.	The process in which ionic compounds when fused or dissolved in water split up into charged particles called ions.	The process in which electricity is used to carry out a non-spontaneous reaction is called electrolysis.



2. Electrodes are not needed. 3. Electricity is not needed. 4. Since there are no electrodes, therefore, ions do not move towards electrodes. 5. After ionization, ions are not discharged. Example: $\text{NaCl}_{(s)} \rightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}$	2. Electrodes are required. 3. Electricity is required. 4. Ions move towards their respective electrodes. 5. Ions are discharged at electrodes to give neutral products. Example: $\text{NaCl}_{(l)} \rightarrow \text{Na}_{(s)} + \text{Cl}_{2(g)}$ At cathode      At anode
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59. Differentiate between electrolytic cell and voltaic cell.

Ans:	Electrolytic cell	Voltaic cell
	1. The electrochemical cell in which electrical energy is converted into chemical energy is called Electrolytic cell. 2. In this cell, non-spontaneous reaction occurs. 3. Electric current is used to drive the non-spontaneous oxidation reduction reaction. 4. Electrolysis takes place in this cell. Example: Down's cell, Nelson's cell	1. The electrochemical cell in which chemical energy is converted into electrical energy is called Voltaic cell. 2. In this cell, spontaneous reaction occurs. 3. Electric current is produced due to spontaneous reaction. 4. Electric conduction takes place in this cell. Example: Daniel's cell, Fuel cells

60. Define Oxidation and Oxidation Number.

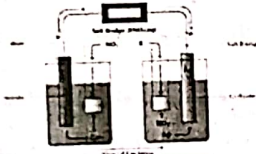
Ans: Oxidation: Removal of electrons is called oxidation.

Oxidation Number: It is the apparent charge on an atom of an element in a compound or a radical. The oxidation number of hydrogen in all its compounds except metal hydrides is +1. In metal hydrides it is -1,  $\text{NaH}$ ,  $\text{Mg}^{2+}\text{H}_2^{(-)}$ .

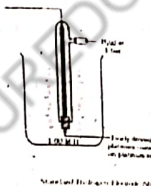
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61. How is voltaic cell represented.

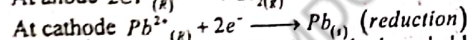
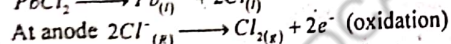
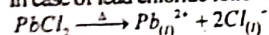
Ans: A Voltaic or galvanic cell is represented as



62. Write down construction of standard hydrogen electrode (SHE).

Ans: It consists of a piece of Pt foil, which is coated with finely divided platinum black, and suspended in one molar solution of  $\text{HCl}$ . Pure hydrogen  $\text{H}_2$  at 1 atm pressure is continuously bubbled in the 1M  $\text{HCl}$  solution. The potential of this electrode is taken arbitrarily zero.63. Explain electrolysis of fused  $\text{PbCl}_2$ .

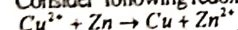
Ans: In case of lead chloride following electrolytic reactions occur at anode and cathode.



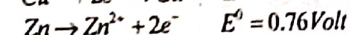
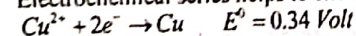
Thus metallic lead is obtained at cathode and chlorine gas at anode.

64. How electrochemical series helps to predict feasibility of a chemical reaction.

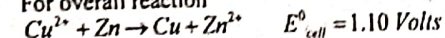
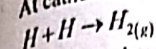
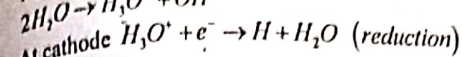
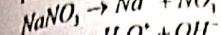
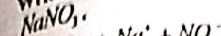
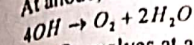
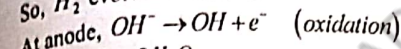
Ans: Consider following redox reaction.



Electrochemical series helps to check, whether this reaction happens or not.



For overall reaction

The overall positive value for reduction potential shows that reaction is possible. Write down reactions at anode and cathode during electrolysis of aqueous solution of  $\text{NaNO}_3$ .So,  $\text{H}_2$  evolves at cathode.So,  $\text{O}_2$  evolves at anode $\text{Na}^+$  and  $\text{NO}_3^-$  ions remain in aqueous solution.

## CHAPTER-10 (ELECTROCHEMISTRY) LONG QUESTIONS (ACCORDING TO ALP SMART SYLLABUS-2020)

State rules for assigning oxidation number of elements with examples? (3 times)

Describe the electrolysis of fused salts and aqueous solution of salts?

Define electrode potential. How electrode potential of Zn is measured?

Write construction and working of voltaic cell? (4 times)

Describe the electrolysis of molten sodium chloride and a concentrated solution of sodium chloride? (4 times)

Define electrolysis. Explain the electrolysis of very dilute solution of  $\text{NaNO}_3$ . (2 times)

What are electrolytic cells? Explain with diagram and give an example of electrolysis of fused salt?

Give four industrial importance of electrolysis process in detail? (4 times)

How can you measure electrode potential of an element with the help of standard hydrogen electrode (SHE)? (2 times)

What is standard hydrogen electrode (SHE)? How is it used to measure the electrode potential of Zinc?

Define standard hydrogen electrode (SHE)? How is it used to measure the electrode potential of copper?

Describe a galvanic cell and explain the functions of salt bridge?

What is electrochemical series? Give its four applications? (8 times)

## Solved Exercise Chapter # 10 ALP SMART SYLLABUS 2020

Q1. Choose the one you consider correct.

(C.W)

(i) The cathodic reaction in the electrolysis of dil.  $\text{H}_2\text{SO}_4$  with Pt electrodes is:

(a) Reduction

(b) Oxidation

(c) Both oxidation and reduction

(d) Neither oxidation or reduction

(ii) Which of the following statements is not correct about galvanic cell?

(a) Anode is negatively charged

(b) Reduction occurs at anode

(c) Cathode is positively charged

(d) Reduction occurs at cathode

(iii) Stronger the oxidizing agent, greater is the:

(a) oxidation potential

(b) reduction potential

(c) redox potential

(d) E.M.F of cell

(iv) If the salt bridge is not used between two half cells, then the voltage.

(a) Decrease rapidly

(b) Decrease slowly

(c) Does not change

(d) Drops to zero

(v) If a strip of Cu metal is placed in a solution of  $\text{FeSO}_4$ :



(a) Cu will be deposited  
(c) Cu and Fe both dissolve

(b) Fe is precipitated out  
(d) No reaction take place

(C.W)

Q.2 Fill in the blank.

(i) The oxidation number of O atom is +2 in  $\text{OF}_2$  and is -1 in  $\text{H}_2\text{O}_2$ .(ii) Conductivity of metallic conductors is due to the low of electrons while that of electrolytes is due to flow of ions.(iii) Reaction taking place at the anode is termed as oxidation and at the cathode is called as reduction.(iv) Equilibrium is set up when a metal is dipped in its own ions.(v) The reduction potential of Zn is -0.76 volts and its oxidation potential is +0.76 volts.

Q.3 Mark the following statements true or false.

(C.W)

(i) In electrolytic conduction, electrons flow through the electrolyte.

(False)

(ii) In the process of electrolysis, the electrons in the external circuit flow from cathode to anode.

(False)

(iii) Sugar is a non-electrolyte in solid form and when dissolved in water will allow the passage of an electric current.

(False)

(iv) A metal will only allow the passage of an electric current when it is in cold state.

(False)

(v) The electrolytic products of aqueous copper (II) chloride solution are copper and chlorine.

(True)

(vi) Zinc can displace iron from its solution.

(True)

(vii) S.H.E. acts as cathode when connected with Cu-electrode.

(False)

(viii) A voltaic cell produces electrical energy at the expense of chemical energy.

(True)

Q.4(c) Calculate the oxidation number of chromium in the following compounds. (C.W)

(i)  $\text{CrCl}_3$  (ii)  $\text{Cr}_2(\text{SO}_4)_3$  (iii)  $\text{K}_2\text{Cr}_2\text{O}_7$  (iv)  $\text{K}_2\text{Cr}_2\text{O}_7$  (v)  $\text{CrO}_3$  (vi)  $\text{Cr}_2\text{O}_3$  (vii)  $\text{Cr}_2\text{O}_7^{2-}$ 

$\text{CrCl}_3$   
Oxidation number of Cl = -1  
Oxidation number of Cr = x

For  $\text{CrCl}_3$   
 $x + 3(-1) = 0$   
 $x - 3 = 0$   
 $x = +3$

$\text{K}_2\text{Cr}_2\text{O}_7$   
Oxidation number of K = +1  
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{K}_2\text{Cr}_2\text{O}_7$   
 $2(1) + 2x + 7(-2) = 0$   
 $2x - 12 = 0$   
 $x = 12/2 = +6$

$\text{Cr}_2\text{O}_3$   
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{Cr}_2\text{O}_3$   
 $2x + 3(-2) = 0$   
 $2x - 6 = 0$   
 $x = 6/2 = +3$

$\text{Cr}_2(\text{SO}_4)_3$   
Oxidation number of S = +6  
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{Cr}_2(\text{SO}_4)_3$   
 $2x + 3[(+6) + 4(-2)] = 0$   
 $2x - 6 = 0$   
 $x = 6/2 = +3$

(d) Calculate the oxidation numbers of the elements underlined in the following compounds.

$\text{K}_2\text{CrO}_4$   
Oxidation number of Cl = +1  
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{K}_2\text{CrO}_4$   
 $2(1) + x + 4(-2) = 0$   
 $x - 6 = 0$   
 $x = +6$

$\text{CrO}_3$   
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{CrO}_3$   
 $x + 3(-2) = 0$   
 $x - 6 = 0$   
 $x = +6$

$\text{Cr}_2\text{O}_7^{2-}$   
Oxidation number of O = -2  
Oxidation number of Cr = x

For  $\text{Cr}_2\text{O}_7^{2-}$   
 $2x + 7(-2) = -2$   
 $2x - 14 = -2$   
 $2x = -2 + 14 = 12$   
 $x = 12/2 = +6$

(i)  $\text{Ca}(\text{ClO}_3)_2$  (ii)  $\text{Na}_2\text{CO}_3$  (iii)  $\text{Na}_3\text{PO}_4$  (iv)  $\text{HNO}_3$ (v)  $\text{Cr}_2(\text{SO}_4)_3$  (vi)  $\text{HPO}_3$  (vii)  $\text{K}_2\text{MnO}_4$ 

$\text{Ca}(\text{ClO}_3)_2$   
Oxidation number of Ca = +2  
Oxidation number of O = -2  
Oxidation number of Cl = x

For  $\text{Ca}(\text{ClO}_3)_2$   
 $(+2) + 2[x + 3(-2)] = 0$   
 $2 + 2x - 12 = 0$   
 $x = 10/2 = +5$

$\text{Na}_2\text{CO}_3$   
Oxidation number of Na = +1  
Oxidation number of O = -2  
Oxidation number of C = x

For  $\text{Na}_2\text{CO}_3$   
 $2(+1) + x + 3(-2) = 0$   
 $x - 4 = 0$   
 $x = +4$

$\text{Cr}_2(\text{SO}_4)_3$   
Oxidation number of Cr = +3  
Oxidation number of O = -2  
Oxidation number of S = x

For  $\text{Cr}_2(\text{SO}_4)_3$   
 $2(+3) + 3[(x) + 4(-2)] = 0$   
 $3x - 18 = 0$   
 $x = 18/3 = +6$

$\text{K}_2\text{MnO}_4$   
Oxidation number of K = +1  
Oxidation number of O = -2  
Oxidation number of Mn = x

For  $\text{K}_2\text{MnO}_4$   
 $(+1) + x + 3(-2) = 0$   
 $x - 5 = 0$   
 $x = +5$

Q.7 Describe the electrolysis of molten sodium chloride, and a concentrated solution of sodium chloride. (H.W)

Consult Text Book Page No. 292-293

Q.8 What is the difference between single electrode potential and standard electrode potential? How can it be measured? Give its importance. (H.W)

Consult Text Book Page No. 295-297

Q.9 Outline the important applications of electrolysis. Write the electrochemical reactions involved therein. Discuss the electrolysis of  $\text{CuSO}_4$  using Cu-electrodes and  $\text{AgNO}_3$  solution using Ag electrode. (H.W)

Consult Text Book Page No. 292-293

Q.10 Describe the construction and working of standard hydrogen electrode. (H.W)

Consult Text Book Page No. 296

Q.11 Is the reaction  $\text{Fe}^{3+} + \text{Ag} \longrightarrow \text{Ag}^+ + \text{Fe}^{2+}$  spontaneous? If not, write spontaneous reaction involving these species. (H.W)Ans:  $\text{Fe}^{3+} + \text{Ag} \longrightarrow \text{Fe}^{2+} + \text{Ag}^+$ In this reaction, Fe is reduced while Ag is oxidized. Therefore  $\text{Fe}^{3+}$  will be act as cathode while  $\text{Ag}^+$  as anode.

Thus, emf of the cell will be

$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$   
 $E^\circ_{\text{cell}} = -0.7994 + (-0.44)$   
 $E^\circ_{\text{cell}} = -0.7994 - 0.44$   
 $E^\circ_{\text{cell}} = -1.2394$

Since emf of cell is negative therefore the cell reaction is non-spontaneous. But if the electrodes are reversed, the cell-reaction becomes spontaneous. i.e.

 $\text{Fe}^{2+} + \text{Ag}^+ \longrightarrow \text{Fe}^{3+} + \text{Ag}^0$  $\text{Na}_2\text{CO}_3$ 

Oxidation number of O = -2  
Oxidation number of Na = +1  
Oxidation number of C = x

For  $\text{Na}_2\text{CO}_3$   
 $2(+1) + x + 3(-2) = 0$   
 $x - 4 = 0$   
 $x = +4$

$\text{HNO}_3$   
Oxidation number of H = +1  
Oxidation number of O = -2  
Oxidation number of N = x

For  $\text{HNO}_3$   
 $+1 + x + 3(-2) = 0$   
 $x - 5 = 0$   
 $x = +5$

$\text{HPO}_3$   
Oxidation number of O = -2  
Oxidation number of H = +1  
Oxidation number of P = x

For  $\text{HPO}_3$   
 $(+1) + x + 3(-2) = 0$   
 $x - 5 = 0$   
 $x = +5$



Q.12 Explain the difference between (a) Ionization and electrolysis.

IONIZATION	ELECTROLYTES
The process in which ionic compounds when fused or dissolved in water split up into charged particles called ions.	The process in which electricity is used to carry out a non-spontaneous reaction is called electrolysis.
Electrodes are not needed.	Electrodes are required.
Electricity is not needed.	Electricity is required.
Since there are no electrodes, therefore ions do not move towards electrodes.	Ions move towards their respective electrodes.
After ionization, ions are not discharged.	Ions are discharged at electrodes to give neutral products.

(b) Electrolytic cell and voltaic cell

Electrolytic Cell	Voltaic cell
1. The electrochemical cell in which electrical energy is converted into chemical energy is called Electrolytic cell. 2. In this cell, non-spontaneous reaction occurs. 3. Electric current is used to drive the non-spontaneous oxidation reduction reaction. 4. Electrolysis takes place in this cell. Example: Down's cell, Nelson's cell	1. The electrochemical cell in which chemical energy is converted into electrical energy is called Voltaic cell. 2. In this cell, spontaneous reaction occurs. 3. Electric current is produced due to spontaneous reaction. 4. Electric conduction takes place in this cell. Example: Daniel's cell, Fuel cells

(c) Conduction through metals and molten electrolytes.

CONDUCTION THROUGH METALS	CONDUCTION THROUGH MOLTEN ELECTROLYTE
Electrical conduction takes place due to free electrons.	Electrical conduction takes place due to ions.
There is no need to convert metal into molten state.	Electrolytes must be converted into molten state for electrical conduction.
In this case, conductance decreases with increase in temperature.	In this case, conductance increases with increase in temperature.
No chemical reaction occurs during conduction.	chemical reaction occurs during conduction.
Chemical composition of metal is not changed during conduction and no new substances are produced.	Since chemical reaction occurs, therefore new substances are produced.
Example: All metals are conductors e.g. Fe, Pb, etc.	Example: Molten Salts e.g. NaCl or their aqueous solutions, acids, bases etc.

Q.13 Describe a galvanic cell explaining the functions of electrodes and the salt bridge.

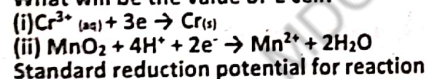
Ans: Consult Text Book Page No. 293-295

Q.14 Write comprehensive notes on:

- (a) Spontaneity of oxidation reduction reactions.  
(b) Electrolytic conduction.

Ans: Consult Text Book Page No. 289-290

Q.15 Will the reaction be spontaneous for the following set of half reactions. What will be the value of E cell? (C.W)



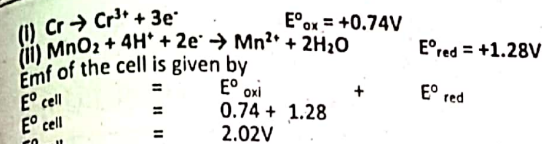
Ans: (i) = -0.74V and for the reaction (ii) +1.28 V

In reaction (i) Cr is reduced from +3 to 0

In reaction (ii) Mn is also reduced from +4 to +2

Both are reduction reactions. Hence these reaction are not possible in these forms.

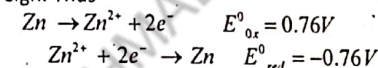
However if reaction (i) is reversed so that Cr is oxidized then the reaction becomes spontaneous and its emf can be calculated as:



Explain the following with reasons.

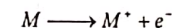
(C.W) The standard oxidation potential of Zn is 0.76 V and its reduction potential - 0.76 V.

Ans: According to the law of conservation of energy, energy can neither be created nor destroyed. Therefore, if standard oxidation potential of Zn is 0.76 V, then its potential for reverse process, i.e. standard reduction potential will also be same but with negative sign. Thus

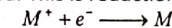


(d) The equilibrium is set up between metal atoms of electrode and ions of metal in a cell.

Ans: When a metal electrode is dipped into the solution of its own ion. There may be two tendencies. Metal atoms from electrode leaves the electron on metal and goes into solution. This is oxidation process.



Metal ions in solution may take up electrons from the metal electrode and deposit as atom on electrode. This is reduction process.



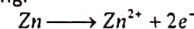
At last, a dynamic equilibrium is established due to same rate of two processes. Thus no further potential difference is developed.

A salt bridge maintains the electrical neutrality in the cell.

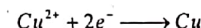
(e) Two half cells are electricity connected by a salt bridge.

Ans: Consider a Zn-Cu cell

During reactions of this cell, Zn half cell continuously lose electrons. Thus in this positive charge is increasing.



While Cu half cell continuously receive electrons, thus it goes on collecting negative charge.



Collection of positive charge in Zn electrode half cell and collection of negative charge in Cu half cell would stop the reaction.

Salt bridge prevents the net accumulation of charges in either beaker. Thus from negative Cu half cell. Negative ions diffuse through the salt bridge into the positive Zn half cell. In this way, salt bridge maintains the two solutions, electricity neutral.

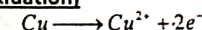
Impure Cu can be purified by electrolytic process.

(g) Ans: Impure Cu can be made pure in an electrolytic cell. Thick sheets of impure copper are made anode, while thin sheets of pure copper are made cathode in the cell. These sheets are placed in an electrolytic solution of  $\text{CuSO}_4$ .

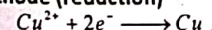
When current is passed through the cell, Cu from anode is oxidized to  $\text{Cu}^{2+}$  ions, which go into the solution. From the solution,  $\text{Cu}^{2+}$  ions are reduced to metallic Cu and deposits as pure Cu on cathode. In this way, impure sheets of Cu (anode) become thin, while pure sheets of pure Cu (cathode) become thick.

The reactions in the cell are

At Anode (oxidation)



At Cathode (reduction)



Thus there is no net reaction in the cell. However, the net result is the purification of Cu.



## CHAPTER-11 (REACTION KINETICS)

### OBJECTIVES (MCQ'S)

(ACCORDING TO ALP SMART SYLLABUS-2020)

- The order of reaction for reaction  $\text{NO} + \text{O}_3 \rightarrow \text{NO}_2 + \text{O}_2$ : (3 times)  
(a) Two (b) Three (c) One (d) Zero
- Which of the following orders, the units of the rate constant are same as that of rate of reaction? (2 times)  
(a) Zero (b) First (c) Second (d) Third
- The unit of the rate constant for zero order reaction is:  
(a)  $\text{mol} \cdot \text{dm}^{-3} \cdot \text{s}^{-1}$  (b)  $\text{mol}^{-1} \cdot \text{dm}^3 \cdot \text{s}^{-1}$  (c)  $\text{dm}^3 \cdot \text{s}^{-1}$  (d)  $\text{mol} \cdot \text{s}^{-1}$
- The unit of the rate constant is the same as that of the rate of reaction in: (3 times)  
(a) first order reaction (b) second order reaction (c) Zero order reaction (d) third order reaction
- The energy of activated complex is:  
(a) Greater than the reactants & products (b) Less than the reactants & products  
(c) Equal to products (d) Equal to reactants
- Velocity constant is the rate of reaction when the concentration of reactant is:  
(a) Zero (b) Unity (c) Two (d) Three
- Arrhenius equation is represented as:  
(a)  $A = K e^{-E_a/RT}$  (b)  $K = A e^{E_a/RT}$  (c)  $A = K e^{E_a/RT}$  (d)  $K = A e^{-E_a/RT}$
- If rate equation of a reaction  $\text{A} + \text{B} \rightarrow \text{products}$  is rate  $= k[\text{A}]^2[\text{B}]$  and A is present in large excess, then order of reaction is:  
(a) 1 (b) 2 (c) 3 (d) 4
- The rate of reaction-----as the reaction proceeds.  
(a) Increase (b) Decrease (c) Remain the same (d) May decrease or increase
- Specific rate constant is equal to rate of reaction, when concentration of reactant is:  
(a) Zero (b) Four (c) Three (d) Unity
- The order of reaction for the reaction  $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$  is:  
(A) zero order (B) first order (C) second order (D) third order
- Half life of a second order reaction is inversely proportional to:  
(A) initial concentration of reactants (B) final concentration of reactants  
(C) initial concentration of products (D) final concentration of products
- Hydrolysis of tertiary butyl bromide is:  
(A) zero order reaction (B) first order reaction (C) pseudo first order reaction (D) second order reaction

#### Answers

1	2	3	4	5	6	7	8	9	10	11	12	13
A	A	A	C	A	B	D	A	B	D	B	A	c

## CHAPTER-11 (REACTION KINETICS)

### SHORT QUESTIONS

(ACCORDING TO ALP SMART SYLLABUS-2020)

#### 1. Rate of Reaction:

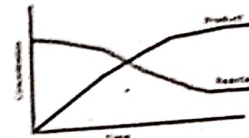
1. Define order of reaction with example? (3 times)

Ans: It is sum of all the exponents of concentrations involved in the rate equation.  
e.g.  $a\text{A} + b\text{B} \rightarrow c\text{C} + d\text{D}$

$$R = k[\text{A}]^x[\text{B}]^y$$

The overall order is  $(a + b)$ .

Rate of chemical reaction is an ever changing (4 times)  
The change in concentration of reactant or product per unit time is called rate of reaction. Rate of reaction never remain constant during different time intervals. Initially rate of reaction is high but decreases with passage of time.  
which is show by diagram.



Differentiate between rate and rate constant of reaction? (2 times)

Rate of reaction	Rate of constant reaction
(i) It is the change in concentration of reactant or product per unit time.	(i) It is the proportionality constant, and is ratio of rate and conc. of reactant.
(ii) $\text{Rate} = \frac{dx}{dt}$	(ii) $K = \frac{\text{Rate}}{[\text{A}]^x[\text{B}]^y}$
(iii) It is an ever-changing parameter. Its unit is $\text{mole} \cdot \text{dm}^{-3} \cdot \text{s}^{-1}$	(iii) It remain constant. Its units are variable.

Define instantaneous and average rate of reaction? (6 times)  
Instantaneous rate of reaction: The rate at any one instant during a specific interval of time is called instantaneous rate of reaction.

Average rate of reaction: The rate of reaction between two specific intervals of time is called average rate of reaction.

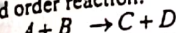
The unit of rate constant of second order reaction in  $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$  but the unit of rate of reaction is  $\text{mol} \cdot \text{dm}^{-3} \cdot \text{s}^{-1}$ . Give reason? (2 times)

Equation for rate of reaction:

$$\text{Rate} = \frac{dx}{dt} = \frac{\text{mol} \cdot \text{dm}^{-3}}{\text{s}}$$

$$\text{Rate} = \text{mol} \cdot \text{dm}^{-3} \cdot \text{s}^{-1}$$

Second order reaction:



$$\text{Rate} = k[\text{A}][\text{B}]$$

$$K = \frac{\text{Rate}}{[\text{A}][\text{B}]} = \frac{\text{mol} \cdot \text{dm}^{-3} \cdot \text{s}^{-1}}{\text{mol} \cdot \text{dm}^{-3} \times \text{mol} \cdot \text{dm}^{-3}}$$

$$K = \text{dm}^3 \cdot \text{mol}^{-1} \cdot \text{s}^{-1}$$

Differentiate between fast step and rate determining step with example? (2 times)

	Fast step	Rate determining step
(i)	A step in a multi-step reaction that has high speed and do not affect rate of reaction is called fast step. $\text{NO}_3 + \text{NO} \xrightarrow{\text{fast}} \text{NO}_2 + \text{NO}_2$	(i) A step in a multi-step reaction that is slowest one and controls the overall rate of reaction is called rate determining step. $\text{NO}_2 + \text{NO}_2 \xrightarrow{\text{slow}} \text{NO}_3 + \text{NO}$

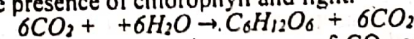
What is pseudo first order reaction? Give an example. (4 times)

Ans: The rate of reaction remains effectively independent of the concentration of water because being a solvent it is present in very large excess. Such type of reactions has been named as pseudo-first order reaction.  
e.g. Hydrolysis of tertiary butyl bromide.

What is zero order reaction? Give one example? (2 times)

Ans: The reaction in which the rate of reaction is independent of the concentration of reacting substances.

All photochemical reactions are zero order reactions e.g. photosynthesis reaction which occurs in the presence of chlorophyll and light.



Rate does not change by changing the amount of  $\text{CO}_2$  or water, however their presence is necessary.

Discuss effect of nature of reactants on rate of reactions? (3 times)

Ans: The rate of reaction depends upon the nature of reacting substances. The chemical



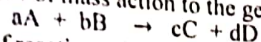
reactivity of the substances is controlled by the electronic arrangements in their outermost orbitals. The elements of I-A group have one electron in their outermost s-orbitals. They react with water more swiftly than those of II-A group elements having two electrons in their outermost s-orbital. Similarly, the neutralization and double decomposition reactions are very fast as compared to those reactions in which bonds are rearranged. Oxidation reduction reactions involve the transfer of electrons and are slower than ionic reactions.

10. **State rate of chemical reaction and give its units?** (3 times)  
 Ans: The rate of reaction is defined as the change in concentration of a reactant or a product divided by the time taken for the change.

The rate of reaction has the units of concentration divided by time. Usually the concentration is expressed in moles  $\text{dm}^{-3}$  and the time in second, thus the units for the reaction rates are moles  $\text{dm}^{-3} \text{s}^{-1}$ .

11. **What is specific rate constant or velocity constant?** (6 times)  
 Ans: The specific rate constant or velocity constant is defined as, the rate of reaction is proportional to the active mass of the reactant or to the product of active masses if more than one reactants are involved in a chemical reaction.

For dilute solutions, active mass is considered as equal to concentration. By applying the law of mass action to the general reaction.



$$\text{Rate of reaction} = k[A]^a[B]^b$$

$$\text{If } [A] = 1 \text{ mol dm}^{-3}, [B] = 1 \text{ mol dm}^{-3}$$

$$\text{Rate of reaction} = k[1]^a[1]^b = k$$

Hence the specific rate constant of a chemical reaction is the rate of reaction when the concentrations of the reaction are unity.

12. **How rate of reaction is determined by electrical conductivity method?**  
 Ans: The rate of a reaction involving ions can be studied by electrical conductivity method. The conductivity of such a solution depends upon the rate of change of concentration of the reacting ions or the ions formed during the reaction. The conductivity will be proportional to the rate of change in the concentration of such ions.

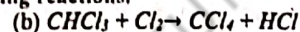
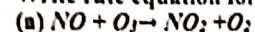
13. **How higher temperature increases the rate of reaction?**  
 Ans: When we increase the temperature the average energy of the molecules increases. The number of those molecules also increases which can form an activated complex after collision. So by increasing temperature, number of effective collision increased and rate of reaction also increases.

2. **Order of Reaction:**

14. **The radioactive decay is always first order reaction. Explain? (10 times)**  
 Ans: The rate of radioactive decay depends on the amount of radioactive substance. Since only one substance is involved in this process therefore it is always a first order reaction.

e.g. Disintegration of  $^{238}\text{U}$  is independent of amount of that substance.

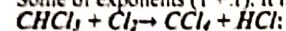
15. **Write rate equation for the following reactions.**



Ans:  $\text{NO} + \text{O}_3 \rightarrow \text{NO}_2 + \text{O}_2$

$$\text{Rate} = k[\text{NO}][\text{O}_3]$$

Some of exponents (1 + 1). It is second order.



$$\text{Rate} = k[\text{CHCl}_3][\text{Cl}_2]^{1/2}$$

The sum of exponents is  $(1 + \frac{1}{2}) = 1.5$ . So order is 1.5.

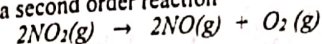
**Compare order of reaction and molecularity?**

Ans:	Order	Molecularity
(i)	The number of molecules whose concentration change in a chemical reaction	(i) The number of molecules involved in a chemical reaction.
(ii)	It is an experimental quantity.	(ii) It is a theoretical quantity.
(iii)	It may be fractional.	(iii) It always whole number.
(iv)	It may be zero.	(iv) It is never zero.

17. **Define with example 2<sup>nd</sup> order reaction?**

Ans: A second order reaction depends on the concentrations of one second order reactant, or two first order reactants.

The reaction rate is given by  $r = k[A]^2$  or by  $r = k[A][B]$   
 For a reaction of type  $2A \rightarrow P$  (where P is the product or products).  
 Examples of a second order reaction



**Half Life:**

**Define half life period. How is it used to determine the order of reaction? (6 times)**  
 Half life period of a reaction is the time required to convert 50% of the reactants into products. For example, the half life period for the decomposition of  $\text{N}_2\text{O}_5$  at  $45^\circ\text{C}$  is 24 minutes.

This method is used to find the order of the reaction as follows:

First order reaction:  $[t_{1/2}]_1 \propto 1/a^0$ , since  $[t_{1/2}]_1 = 0.693/k$

Second order reaction:  $[t_{1/2}]_2 \propto 1/a^1$ , since  $[t_{1/2}]_2 = 1/ka$

Third order reaction:  $[t_{1/2}]_3 \propto 1/a^2$ , since  $[t_{1/2}]_3 = 1.5/ka^2$

Where  $[t_{1/2}]_1$ ,  $[t_{1/2}]_2$  and  $[t_{1/2}]_3$  are the half life periods for 1<sup>st</sup>, 2<sup>nd</sup> and 3<sup>rd</sup> order reactions respectively and 'a' is the initial concentration of reactions.

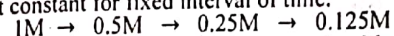
In general for the reaction of nth order.  $t_{1/2} \propto 1/a^{n-1}$

The half life period of any order is thus inversely proportional to the initial concentration raised to the power of one less than the order of that reaction.

50% of hypothetical first order reaction completes in one hour. The remaining

50% needs more than one hour to convert itself into products why?

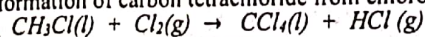
First order reaction is independent of initial concentration of reactant. Suppose initial amount of reactant is  $1 \text{ mol dm}^{-3}$ . One half of the initial value decreases in half time ( $t_{1/2}$ ). After half time, initial value decrease to  $0.5 \text{ mol dm}^{-3}$ . For the next half time, it decreases upto only  $0.25 \text{ mol dm}^{-3}$ . Decreases in concentration of reactant is not constant for fixed interval of time.



The order of a reaction may be in fractions. Justify with the help of an example?

The order of the reaction is usually positive integer or a zero, but it can also be a fraction or can have a negative value.

Consider the formation of carbon tetrachloride from chloroform.



$$\text{Rate} = k[\text{CHCl}_3][\text{Cl}_2]^{1/2}$$

The sum of exponents will be  $1 + 1/2 = 1.5$ , so the order of this reaction is 1.5.

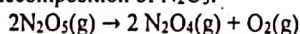
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21. **The sum of the co-efficients of a balanced chemical equation is not necessarily important to give the order of reaction. Give reasons in support of your answer.**

Ans: Rate of reaction is an experimentally determined quantity. It cannot be predicated from balanced chemical equation. The order of reaction is determined from experimentally obtained rate equation. Through balanced chemical equation we can only calculate the molecularity of reaction.

**Example:**

Consider the decomposition of  $\text{N}_2\text{O}_5$ .



Apparently it seems to be 2<sup>nd</sup> order reaction but experimentally determined rate equation is:

$$\text{Rate} = k[\text{N}_2\text{O}_5]$$

It indicates that it is first order reaction.

**2019**

22. **The order of reaction may be in fraction. Explain.**

Ans: Consider the following reaction  $\text{CHCl}_3 + \text{Cl}_2 \rightarrow \text{CCl}_4 + \text{HCl}$

$$\text{Rate} = k[\text{CHCl}_3][\text{Cl}_2]^{1/2}$$

The sum of exponents will be  $1 + 1/2 = 1.5$  so it is clear that order of reaction may also be in fractions.



**CHAPTER-11 (REACTION KINETICS)**  
**(LONG QUESTIONS)**  
**(ACCORDING TO ALP SMART SYLLABUS-2020)**

- Define these terms: (i) Rate of reaction (ii) Order of reaction?
- Name various factors affecting rate of reactions. Explain any one in detail? (2 times)
- Discuss different factors which affect the rate of reactions? (2 times)
- Explain effect of concentration of reaction on rate of reaction. (10 times)
- How light and surface area affect the rate of reactions? (9 times)
- Define catalysis. Give its types and describe one type with an example?
- How does Arrhenius equation help us to calculate the energy of activation of a reaction?
- Define half-life period and order of reaction. Describe half-life method to determine the order of reaction?

**Solved Exercise Chapter # 11**  
**ALP SMART SYLLABUS 2020**

**Q.1 Multiple choice questions.**

- In zero order reaction, the rate is independent of  
 a) temperature of reaction. (b) concentration of reactants,  
 c) concentration of products (d) none of these
- If the rate equation of a reaction  $2A + B \rightarrow \text{products}$  is, rate  $= k[A]^2[B]$ , and A is present in large excess, then order of reaction is  
 a) 1 (b) 2 (c) 3 (d) none of these
- The rate of reaction  
 a) increases as the reaction proceeds.  
 b) decreases as the reaction proceeds.  
 c) remains the same as the reaction proceeds.  
 d) may decrease or increase as the reaction proceeds.
- With increase of  $10^\circ\text{C}$  temperature the rate of reaction doubles. This increase in rate of reaction is due to:  
 a) decrease in activation energy of reaction.  
 b) decrease in the number of collisions between reactant molecules.  
 c) increase in activation energy of reactants.  
 d) increase in number of effective collisions.
- The unit of the rate constant is the same as that of the rate of reaction in  
 (a) first order reaction. (b) second order reaction.  
 (c) zero order reaction. (d) third order reaction.

**Q.2 Fill in the blanks with suitable words.**

- The rate of an endothermic reaction increases with the increase in temperature.
- All radioactive disintegration nuclear reactions are of first order
- For a fast reaction the rate constant is relatively greater and half-life is short.
- The second-order reaction becomes first order if one of the reactants is in large excess.

Arrhenius equation can be used to find out energy of activation of a reaction.

Indicate true or false as the case may be.

- The half life of a first order reaction increases with temperature. (False)
- The reactions having zero activation energies are instantaneous. (True)
- There is difference between rate law and the law of mass action. (False)
- The order of reaction is strictly determined by the stoichiometry of the balanced equation. (False)

11<sup>th</sup> Class

149

A Plus Chemistry (ALP Smart Syllabus-2020)

**Q4. What is chemical kinetics? How do you compare chemical kinetics with chemical equilibrium and thermo dynamics. (H.W)**

**Ans: Chemical kinetics**

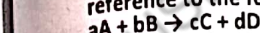
- It is the study of rates of reactions and the factors, which effects these rates like concentration, temperature, catalyst etc.
- It gives information about mechanism of reactions.
- Both reversible and irreversible reactions are studied in chemical kinetics.

**Chemical equilibrium**

- It is the state of a reversible reaction at which the rate of forward reaction becomes equal to the rate of reverse reaction.
- It is only concerned with reversible reactions.
- Le-chatlier's principle is used to study the effects temperature, pressure and concentration on equilibrium state.

**Thermodynamics**

- It is the study of energy changes.
- It tells about the feasibility of a chemical reaction.
- The rate of a chemical reaction with respect to products is written with positive sign, but with respect to reactants is written with a negative sign. Explain it with reference to the following hypothetical reaction. (H.W)



- Ans:** During a chemical reaction, reactants are converted into products and their concentration change. Thus rate of a reaction can be expressed with respect to any reactant or product.  
 However as the reaction proceeds, the concentration of reactants decreases with time and the concentration of products increases with time. Thus, rate of chemical reaction with respect to products is written with positive sign but with respect to reactants, it is written with negative sign.

Hence rate of above hypothetical reaction can be expressed as

$$\text{Rate} = -\frac{1}{a} \frac{d[A]}{dt} = -\frac{1}{b} \frac{d[B]}{dt} = \frac{1}{c} \frac{d[C]}{dt} = \frac{1}{d} \frac{d[D]}{dt}$$

**Q6. What are instantaneous and average rates? Is it true that the instantaneous rate of a reaction at the beginning of the reaction is greater than average rate and becomes far less than the average rate near the completion of reaction? (H.W)**

- Ans:** The rate of reaction between specific time intervals is called the average rate of reaction. While the rate at any one instant during the interval is called instantaneous rate.

**Q7. Differentiate between Rate and rate constant of a reaction (H.W)**

Rate of a reaction	Rate constant of a reaction
It is the change in concentration of reactants or products divided by the time taken for the change.	It is the rate of a reaction when the concentration of reactants are unity.
It changes with time.	It does not changes with time.
It depends upon the concentration of reactants	It is independent of the Concentration of reactants
It is a variable quantity.	It is a constant quantity.
Its units are $\text{mol dm}^{-3}\text{s}^{-1}$	Its unit depend upon the order of reaction.
Consider a general reaction $A + B \rightarrow C + D$ Its rate is given as $\text{Rate} = k[C][D]$ , where k is the rate of constant.	

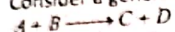
**Q8. Justify the following statements (C.W)**

- Rate of chemical reaction is an ever changing parameter under the given conditions.



Ans: According to the law of mass action, rate of a chemical reaction is directly proportional to the concentration of reactants. When the reaction starts, the concentration of reactants is high, therefore, rate of reaction is fast. As the concentration of reactants is decreased, the rate of a reaction is also decreased. At the end of reaction, the reaction becomes very slow. Hence rate of a chemical reaction is an ever changing parameter.

(iii) The reaction rate decreases every moment but rate constant 'k' of the reaction is a constant quantity, under the given conditions. According to the law of mass action, rate of a chemical reaction is directly proportional to the concentration of reactants. Consider a general reaction.



Its rate is given as:

Rate =  $k[A][B]$ , where k is the rate of constant.

As the reaction proceeds, concentration of reactants is decreased, therefore, rate is also decreased. However rate constant k is not changed. It is proportionality constant.

(iii) 50% of a hypothetical first order reaction completes in one hour. The remaining 50% needs more than one hour to complete.

Ans: The time required for half of the reactants (i.e. 50%) to convert into products is called half life time of the reaction. For a first order reaction, half life time is a constant quantity under given conditions.

For the above hypothetical first order reaction, half life time is one hour. Therefore if we start with 100% reactants then the reaction will proceed as

Time(hour)	0	1	2	3	4
Amount of Reactant	100	50	25	12.5	6.75

Hence, after every one hour, amount of reactant becomes half.

It shows that 50% of the reactants are consumed. Thus in a total of two hours, 75% of the total amount is consumed.

Hence 50% reactants are consumed in one hour but the remaining 50% will take many hours to be consumed.

(iv) The radioactive decay is always a first order reaction.

Ans: The rate of radioactive decay depends on the amount of radioactive substance. Since only one substance (reactant) is involved in this process, therefore, it is always a first order reaction.

(v) The unit of rate constant of a second order reaction is  $\text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$ , but the unit of rate of reaction is  $\text{mol dm}^{-3} \text{s}^{-1}$ .

Ans: Rate of reaction is the change in concentration with change in time. Therefore its units are given as:

$$\text{Rate} = \frac{\Delta c}{\Delta t} = \frac{\text{mol / dm}^3}{\text{sec}} = \text{mol dm}^{-3} \text{sec}^{-1}$$

However for a second order reaction the rate of reaction is directly proportional to the concentration of two reactants e.g. A and B, i.e.

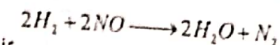
Rate =  $k[A][B]$ , where k is the rate of constant.

Since the units of rate =  $\text{mol dm}^{-3} \text{sec}^{-1}$  and concentration of A and B are expressed in  $\text{mol dm}^{-3}$  therefore units of k will be.

$$k = \frac{\text{Rate}}{[A][B]} = \frac{\text{mol dm}^{-3} \text{sec}^{-1}}{\text{mol dm}^{-3} \times \text{mol dm}^{-3}} = \text{dm}^3 \text{mol}^{-1} \text{s}^{-1}$$

(vi) The sum of the coefficients of a balanced chemical equation is not necessarily important to give the order of a reaction.

Ans: A balanced chemical equation only gives the stoichiometry of the reaction. It does not deal with the mechanism of the reaction. However, order of reaction is determined experimentally and it depends upon the mechanism of reaction. Actually it depends upon the slow step of the reaction. Therefore the co-efficients in balanced chemical equation has no relation with order of reaction. e.g., consider reaction.



Rate of this reaction is

$$\text{Rate} = k[\text{H}_2][\text{NO}]^2$$

Hence its order is  $1 + 2 = 3$  (third order reaction)

However the sum of co-efficients in balanced chemical equation =  $2 + 2 = 4$

Sometimes the sum of co-efficient may be equal to the order of reaction.

(vii) The order of a reaction is obtained from the rate expression of a reaction and the rate expression is obtained from the experiment.

Ans: Consider a general reaction



Rate expression for this reaction will be.

$$\text{Rate} = k[A]^a[B]^b$$

'a' and 'b' indicates the number of molecules involved in the slow step of the reaction. These are determined experimentally. Rate expression is, therefore, an experimental expression. Hence order of reaction (i.e.  $a+b$ ) is also an experimental quantity.

Q9: Explain that half life method for measurement of the order of a reaction can help us to measure the order of even those reactions which have a fractional order. (H.W)

Ans: Using half life method, order of reaction can be determined by the equation.

$$n = 1 + \frac{\log\left(\frac{t_1}{t_2}\right)}{\log\left(\frac{a_2}{a_1}\right)}$$

Where n = order of reaction.

$t_1$  is the half life time when the concentration of reactants is  $a_1$  and  $t_2$  is the half life time when the concentration of reactants is  $a_2$ .

Thus knowing  $t_1$  and  $t_2$  at  $a_1$  and  $a_2$  respectively, order of reaction 'n' can be determined.

The answer can be in whole number or in fraction. Hence, half life time method can be used to determine both simple and fractional order.

Q15. How does Arrhenius equation help us to calculate the energy of activation of a reaction? (H.W)

Ans: Consult Text book Page No. 322-324

Q19. In the reaction of NO and  $\text{H}_2$ , it was observed that equimolecular mixture of gases at 340.5 mm Hg pressure was half changed in 102 seconds. In another experiment with an initial pressure of 288 mm of Hg, the reaction was half completed in 140 seconds. Calculate the order of reaction. (C.W)

Ans: For 1<sup>st</sup> experiment

Initial concentration =  $a_1 = 340.5 \text{ mm}$

Half life time =  $t_1 = 102 \text{ sec}$

For 2<sup>nd</sup> experiment

Initial concentration =  $a_2 = 288 \text{ mm}$

Half life time =  $t_2 = 140 \text{ sec}$

Order of reaction is given by

$$n = 1 + \frac{\log\left(\frac{t_1}{t_2}\right)}{\log\left(\frac{a_2}{a_1}\right)}$$

$$n = 1 + \frac{\log\left(\frac{102}{140}\right)}{\log\left(\frac{288}{340.5}\right)}$$



$$n = 1 + \frac{-0.1375}{-0.0727}$$

$$n = 1 + \frac{0.1375}{0.0727} = 2.89 \approx 3 \quad (\text{C.W})$$

Q20. A study of chemical kinetics of a reaction  
 $A + B \rightarrow \text{Products}$

gave the following data at 25 °C. Calculate the rate law.

Exp No.	(A)	(B)	Rate
1	1	0.15	$4.2 \times 10^{-4}$
2	2	0.15	$8.4 \times 10^{-4}$
3	1	0.2	$5.6 \times 10^{-4}$

Ans: Comparison of Exp. (1) and (2) shows that keeping [B] constant and doubling [A], doubles the reaction rate.  
 Thus Rate  $\propto [A]^1$   
 Comparison of Exp. (1) and (3) shows that keeping [A] constant and changing [B], from 0.15 to 0.2 rate changes from  $4.2 \times 10^{-4}$  to  $5.6 \times 10^{-4}$   
 Thus Rate  $\propto [B]^1$   
 Hence overall order of rate is given by.  
 Rate  $\propto [A]^1[B]^1$   
 Rate =  $k[A]^1[B]^1$

Thus overall order of reaction is  $1 + 1 = 2$   
 Q21. Some reactions taking place around room temperature have activation energies around  $50 \text{ kJ mol}^{-1}$ . (C.W)

(I) What is the value of the factor  $e^{-\frac{E_a}{RT}}$  at 25°C?

Ans:  $E_a = 50 \text{ kJ mol}^{-1} = 50000 \text{ J mol}^{-1}$   
 $T = 25^\circ\text{C} + 273 = 298 \text{ K}$   
 $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

Thus  

$$e^{-\frac{50000}{8.314 \times 298}} = e^{-20.181} = 1.72 \times 10^{-9}$$

(II) Calculate this factor at 35 °C and at 45°C and note the increase in this factor for every 10°C rise in temperature.

Ans:  $E_a = 50 \text{ kJ mol}^{-1} = 50000 \text{ J mol}^{-1}$   
 $T = 35^\circ\text{C} + 273 = 308 \text{ K}$   
 $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

Thus  

$$e^{-\frac{50000}{8.314 \times 308}} = e^{-19.52} = 3.33 \times 10^{-9}$$

At 45°C  
 $E_a = 50 \text{ kJ mol}^{-1} = 50000 \text{ J mol}^{-1}$   
 $T = 45^\circ\text{C} + 273 = 318 \text{ K}$   
 $R = 8.314 \text{ J mol}^{-1} \text{ K}^{-1}$

Thus  

$$e^{-\frac{50000}{8.314 \times 318}} = e^{-18.91} = 6.13 \times 10^{-9}$$

Exponential factor at 25°C =  $1.72 \times 10^{-9}$   
 Exponential factor at 35°C =  $3.33 \times 10^{-9}$   
 Exponential factor at 45°C =  $6.13 \times 10^{-9}$

Hence exponential factor doubles for every 10°C rise in T.

(III) Prove that for every 10°C rise in of temperature, the factor doubles and so rate constant also doubles.

Rate is given by Arrhenius eq.

$$k = Ae^{-\frac{E_a}{RT}}$$

For a reaction A is almost constant

Write calculations of part (ii) here

Result of part (ii) shows that for every 10°C rise in T, exponential factor is double.

Hence rate is doubled for every 10°C rise in T.

## Full Book Model Paper-1

Chemistry

(Inter Part-I Class 11<sup>th</sup>)

Time : 20

Objective

Marks : 17

Note: You have choices for each objective type question as A, B, C and D. The choice which you think is correct, fill that circle in front of that question number with marker or pen. Cutting or filling two or more circles will result in zero mark in that question.

Q.1: Answer all the following Multiple Choice Questions.

- Mass of one mole of electrons is: (16 times)  
 (a) 0.55 mg (b) 0.184 mg (c) 1.673 mg (d) 1.008 mg
- The number of moles of  $\text{CO}_2$  which contains 8.0g of oxygen is: (19 times)  
 (a) 0.25 (b) 0.50 (c) 1.0 (d) 1.50
- Iodine dissolve in water in the presence of KI due to formation of which one of the following species: (1 time)  
 (a)  $\text{I}_2$  (b)  $\text{I}^{(-)}$  (c)  $\text{I}_3^{(-)}$  (d)  $\text{I}_4$
- Partial pressure of oxygen in human lungs in torr is: (6 times)  
 (a) 161 (b) 116 (c) 159 (d) 760
- Which of the following liquid has highest boiling point: (2 times)  
 (a) HCl (b) HBr (c)  $\text{H}_2\text{O}$  (d)  $\text{Br}_2$
- $^{65}_{29}\text{Cu} + {}^1_0\text{n} \rightarrow ^{66}_{29}\text{Cu} + \text{X}$  (5 times)  
 (a) Electron (b) Proton (c)  $\beta$ -rays (d)  $\text{h}\nu/\gamma$ -rays
- Octet rule is not obeyed during its formation: (4 times)  
 (a)  $\text{NF}_3$  (b)  $\text{CF}_4$  (c)  $\text{PCl}_5$  (d)  $\text{CCl}_4$
- Which of the following is not a state function: (3 times)  
 (a) Pressure (b) Volume (c) Temperature (d) Heat
- The relationship between  $K_p$  and  $K_c$  is given by: (3 times)  
 (a)  $K_c = K_p(P)^{\Delta n}$  (b)  $K_c = K_p(P/N)^{\Delta n}$  (c)  $K_p = K_c(RT)^{\Delta n}$  (d)  $K_p = K_c(RT)^{-\Delta n}$
- An aqueous solution of ethanol in water may have vapour pressure: (2 times)  
 (A) equal to water (B) more than that of water  
 (C) equal to ethanol (D) less than that of water
- The oxidation number of chromium in  $\text{K}_2\text{Cr}_2\text{O}_7$  is: (4 times)  
 (a) +14 (b) +12 (c) +6 (d) +13
- Which of the following orders, the units of the rate constant are same as that of rate of reaction? (2 times)  
 (a) Zero (b) First (c) Second (d) Third
- The unit of the rate constant is the same as that of the rate of reaction in: (3 times)  
 (a) first order reaction (b) second order reaction (c) Zero order reaction (d) third order reaction
- The largest number of molecules are present in: (27 times)  
 (a) 2.8g of  $\text{CO}$  (b) 3.6g of  $\text{H}_2\text{O}$  (c) 4.6g of  $\text{C}_2\text{H}_5\text{OH}$  (d) 5.4g of  $\text{N}_2\text{O}_5$
- Partial pressure of oxygen in the air is: (2 times)  
 (a) 156 torr (b) 157 torr (c) 158 torr (d) 159 torr
- When  $a=b=c$  and  $\alpha=\beta=\gamma=90^\circ$  then the crystal system is: (2 times)  
 (a) Hexagonal (b) Monoclinic (c) Cubic (d) Tetragonal
- Lyman series lies in: (3 times)  
 (a) U.V region (b) Visible region (c) I.R region (d) Microwave region



# Full Book Model Paper-1

Chemistry

(Inter Part-I Class 11<sup>th</sup>)

Time: 2:40 Hours

Marks: 68

## Subjective

Note: Section I is compulsory, Attempt any 3 questions from Section II.

### Section-I

Q.2: Write short answers to any Eight parts.

- 180g of glucose and 342g of sucrose have the same number of molecules but different number of atoms present in them. Explain? (8 × 2 = 16)  
(5 times)
- Define mole and molar volume? (4 times)
- Calculate the moles of 100g of silicon? (Atomic Mass of Si is 28) (2 times)
- Define sublimation with two examples? (9 times)
- Define sublimation and partition law? (2 times)
- Iodine is more soluble in water in the presence of KI. Discuss? (13 times)
- What is absolute zero? What happens to real gases while approaching this temperature? (11 times)
- State Charles's Law and write its mathematical form? (2 times)
- Intermolecular forces are weaker than intermolecular forces why? (2 times)
- Define dipole-dipole forces with example? (4 times)
- Why HF is weaker acid than HCl? (5 times)
- Describe the importance of vacuum distillation? (8 × 2 = 16)  
(2 times)

Q.3: Write short answers to any Eight parts.

- How will you prove that Cathode Rays travel in Straight Lines? (8 times)
- Write four properties of positive rays? (3 times)
- Compare bond strength of polar and non-polar molecules? (5 times)
- Why atomic radius is greater than cationic radius? (5 times)
- Why cationic radius is smaller than parent atom? (8 times)
- Burning of a candle is a spontaneous process. Justify? (3 times)
- What are thermochemical reaction, give their types? (7 times)
- Describe system and surrounding? (7 times)
- The e/m values of positive rays obtained from hydrogen gas is 1836 times less than that of cathode rays. Justify? (8 times)
- Cathode rays are negatively charged? Explain it with diagram. (10 times)
- Why is it necessary to decrease the pressure in discharge tube to get cathode rays? (2 × 6 = 12)  
(2 times)
- Define octet rule. Give two examples of compounds which do not obey this rule? (2 times)

Q.4: Write short answers to any Six parts.

- Why the atoms of the elements other than noble gases combine with one another? (3 times)
- What is paramagnetic character? Give the reason for paramagnetic character of oxygen? (2 times)
- Is it true that a non-spontaneous process never happen in universe? Explain it. (7 times)
- Define state function? (8 times)
- Difference between reversible and irreversible reactions? (8 times)
- State law of mass action? (2 times)
- Write two applications of equilibrium constant? (3 times)
- Calculate pH of  $10^{-3}$  mol  $\text{dm}^{-3}$  HCl (4 times)
- One molal solution of glucose in water is dilute as compared to one molar solution of glucose. Justify? (3 times)

### Section-II

Note: Attempt any three (3) questions:

- Write a note on Avogadro's number. (2 times)
- State and explain Dalton's Law of Partial pressures. Derive an expression of calculating Partial pressure of a gas. (4 times)
- What is H-bonding? Discuss H-bonding in biological compounds. (3 times)
- Discuss properties of cathode rays? (5 times)
- Give important postulates of VSEPR theory? (6 times)
- How  $\Delta H$  can be determined by using bomb calorimeter? (2 times)
- Define common ion effect. Give its two applications? (2 times)
- Explain positive and negative deviation of non-ideal solution. (3 times)
- State rules for assigning oxidation number of elements with examples? (10 times)
- How does Arrhenius equation help us to calculate the energy of activation of a reaction? (10 times)

## Objective

Note: You have choices for each objective type question as A, B, C and D. The choice which you think is correct, fill that circle in front of that question number with marker or pen. Cutting or filling two or more circles will result in zero mark in that question.

Q.1: Answer all the following Multiple Choice Questions.

- The volume occupied by 1.4g of  $\text{N}_2$  at S.T.P in  $\text{dm}^3$  is: (14 times)  
(a) 1.12  $\text{dm}^3$  (b) 2.24  $\text{dm}^3$  (c) 11.2  $\text{cm}^3$  (d) 22.4  $\text{dm}^3$
- In 1879, plasma was identified by scientist: (1 time)  
(a) John Dalton (b) Chadwick (c) William Crookes (d) Soddy
- Mass of 22.4  $\text{dm}^3$  of  $\text{N}_2$  at STP is: (3 times)  
(a) 28 gm (b) 14 gm (c) 1.4 gm (d) 2.8 gm
- Ice occupies more spaces than liquid water upto: (2 times)  
(a) 9% (b) 10% (c) 11% (d) 12%
- Mass of electron is: (3 times)  
(a)  $9.1095 \times 10^{31}$  kg (b)  $9.1095 \times 10^{-31}$  kg (c)  $9.1095 \times 10^{-27}$  kg (d)  $9.1095 \times 10^{-31}$  g
- The electrons in a sub shell are filled according to formula (2 times)  
(a)  $2n^2$  (b)  $2(2l+1)$  (c)  $(2l+1)$  (d) None of these
- Bond order of  $\text{O}_2^{2-}$  is: (2 times)  
(a) Zero (b) one (c) Two (d) Three
- The pressure of oxygen inside the bomb calorimeter is: (2 times)  
(a) 100 atm (b) 50 atm (c) 25 atm (d) 20 atm
- The value of  $K_w$  at  $25^\circ\text{C}$  is: (3 times)  
(a)  $0.11 \times 10^{-14}$  (b)  $0.30 \times 10^{-14}$  (c)  $1 \times 10^{-14}$  (d)  $3 \times 10^{-14}$
- Molarity of pure water is: (4 times)  
(a) 1 (b) 18 (c) 55.5 (d) 6
- Oxidation number of "Cr" in  $\text{K}_2\text{CrO}_4$  is: (2 times)  
(a) +2 (b) +4 (c) +6 (d) +8
- One mole of  $\text{SO}_2$  contain (12 times)  
(a)  $6.02 \times 10^{23}$  atoms of Oxygen (b)  $18.01 \times 10^{23}$  molecules of  $\text{SO}_2$   
(c)  $6.02 \times 10^{23}$  atoms of Sulphur (d) 4 gram atoms of  $\text{SO}_2$
- The most common solvent used in solvent extraction is:  
(a) Acetone (b) Ethanol (c) Rectified spirit (d) Diethyl ether
- Transition temperature of  $\text{KNO}_3$  is: (2 times)  
(a)  $13.2^\circ\text{C}$  (b)  $95.5^\circ\text{C}$  (c)  $128^\circ\text{C}$  (d)  $32.2^\circ\text{C}$
- When the azimuthal quantum number is 3 then 'm' can have  
(a) 5 values (b) 7 values (c) 2 values (d) 3 values
- Which molecule has  $\text{sp}^2$  hybridization? (2 times)  
(a)  $\text{CH}_4$  (b)  $\text{C}_2\text{H}_4$  (c)  $\text{C}_2\text{H}_2$  (d)  $\text{C}_2\text{H}_6$
- The pH of human blood is: (4 times)  
(a) 8.0 (b) 7.53 (c) 7.63 (d) 7.35



## Chemistry

**Full Book Model Paper-1**(Inter Part-I Class 11<sup>th</sup>)

Subjective

Time: 2:40 Hours  
Marks : 68**Note: Section I is compulsory, Attempt any 3 questions from Section II.****Section-I****Q.2: Write short answers to any Eight parts.**

(8 × 2 = 16)

- i. Mg atom is twice heavier than that of carbon atom. Comment? (8 times)
- ii. What do you know about gram atom? (2 times)
- iii. Calculate number of gram atoms of Na when its mass is 0.1kg? (2 times)
- iv. Give the main characteristics of the solvent used for crystallization? (6 times)
- v. How mixture of  $\text{NH}_4\text{Cl}$  and  $\text{NaCl}$  can be separated? (2 times)
- vi. State distribution law or partition law? (12 times)
- vii. Define Boyle's law and give its mathematical expression? (3 times)
- viii. Convert  $-40^\circ\text{F}$  temperature to (a) Centigrade scale (b) Kelvin scale? (3 times)
- ix. Ice floats on  $\text{H}_2\text{O}$ . Give reason? (6 times)
- x. Water is liquid at room temperature while  $\text{H}_2\text{S}$  is a gas. Comment? (3 times)
- xi. Give two important uses of liquid crystals? (12 times)
- xii. Define liquid crystal with one example. (4 times)

**Q.3: Write short answers to any Eight parts.**

(8 × 2 = 16)

- i. Whatever gas is used in the discharge tube the nature of the cathode rays remains the same, why? (7 times)
  - ii. Justify that  $e/m$  value of positive rays for different gases are different but those for cathode rays, the  $e/m$  values are same. (2 times)
  - iii. The positive rays are also called canal rays. Explain? (7 times)
  - iv. Define ionization energy. Give two factors on which it depends? (4 times)
  - v. Why cationic radii is smaller than anionic radii? (4 times)
  - vi. Why the atomic radii of the atoms cannot be determined precisely? (4 times)
  - vii. Differentiate between internal energy and enthalpy? (7 times)
  - viii. Prove  $\Delta E = q_v$ ? (4 times)
  - ix. State first law of thermodynamics with its mathematical form? (5 times)
  - x. How  $K_c$  predict the extent of chemical reaction? (5 times)
  - xi. Define  $K_f$  and  $K_b$  for reversible reactions? (2 times)
  - xii. What are  $K_c$  and  $K_p$  and how these are related? (2 times)
- Q.4: Write short answers to any Six parts.** (2 × 6 = 12)
- i. What is the effect of catalyst of equilibrium constant? (6 times)
  - ii. The sum of all the mole fractions is equal to one (unity). Discuss? (3 times)
  - iii. Give two definitions of Raoult's law? (4 times)
  - iv. Why the non-ideal solutions do not obey the Raoult's law? (3 times)
  - v. Differentiate between ideals and non-ideal solutions? (9 times)
  - vi. Define electrochemistry? (4 times)
  - vii. Define oxidation state with two examples? (3 times)
  - viii. Calculate the oxidation number of chromium in:  $\text{Cr}_2\text{O}_7^{2-}$ ? (3 times)
  - ix. Differentiate between rate and rate constant of reaction? (2 times)

**Section-II****Note: Attempt any three (3) questions:**

(3 × 8 = 24)

5. (a) A well known ideal gas is enclosed in a container having volume  $500 \text{ cm}^3$  at STP its mass comes out to be  $0.72 \text{g}$ . What is the molar mass of this gas. (3 times)
- (b) Calculate the mass of  $1 \text{ dm}^3$  of  $\text{NH}_3$  gas at  $30^\circ\text{C}$  and  $100 \text{ mm/Hg}$  pressure, considering the  $\text{NH}_3$  is behaving ideally? (6 times)
- (a) Define liquid crystals. Discuss important uses of liquid crystal. (6 times)
- (b) Derive the formula for calculating the energy of an electron in  $n^{\text{th}}$  orbit using Bohr's model? (4 times)
- (a) Explain  $\text{sp}^3$  hybridization with the help of two examples? (4 times)
- (b) State law of thermodynamics and prove that  $\Delta E = q_v$ ? (5 times)
- (a)  $\text{N}_2(\text{g})$  and  $\text{H}_2(\text{g})$  combine to give  $\text{NH}_3(\text{g})$ . The value of  $K_c$  in this reaction at  $500^\circ\text{C}$  is  $6.0 \times 10^{-2}$ . Calculate the value of  $K_p$  for this reaction? (9 times)
- (b) State and explain Raoult's law in three forms. (11 times)
- a) What is electrochemical series? Give its four applications? (8 times)
- b) Discuss different factors which affect the rate of reactions? (2 times)