

**Support Material Preparation Team**  
**Chemistry – XI**

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# Course Structure

## Class : XI (Theory) (2016-17)

### Chemistry

Total period (Theory 160 + Practical 60)

Time : 3 Hours]

Total Marks : 70

Unit No.	Title	No. of Periods	Marks
Unit I	Some Basic Concepts of Chemistry	12	
Unit II	Structure of Atom	14	11
Unit III	Classification of Elements and Periodicity in Properties	08	04
Unit IV	Chemical Bonding and Molecular Structure	14	
Unit V	States of Matter: Gases and Liquids	12	
Unit VI	Chemical Thermodynamics	16	21
Unit VII	Equilibrium	14	
Unit VIII	Redox Reactions	06	
Unit IX	Hydrogen	08	16
Unit X	s -Block Elements	10	
Unit XI	Some p -Block Elements	14	
Unit XII	Organic Chemistry: Some basic Principles and Techniques	14	
Unit XIII	Hydrocarbons	12	18
Unit XIV	Environmental Chemistry	06	
<b>Total</b>		<b>160</b>	<b>70</b>

**Unit I : Some Basic Concepts of Chemistry****12 Periods**

General Introduction: Importance and scope of chemistry.

Nature of matter, laws of chemical combination, Dalton's atomic theory : concept of elements, atoms and molecules.

Atomic and molecular masses, mole concept and molar mass, percentage composition, empirical and molecular formula, chemical reactions, stoichiometry and calculations based on stoichiometry.

**Unit II : Structure of Atom****14 Periods**

Discovery of Electron, Proton and Neutron, atomic number, isotopes and isobars. Thomson's model and its limitations. Rutherford's model and its limitations, Bohr's model and its limitations, concept of shells and subshells, dual nature of matter and light, de Broglie's relationship, Heisenberg uncertainty principle, concept of orbitals, quantum numbers, shapes of *s*, *p* and *d* orbitals, rules for filling electrons in orbitals : Aufbau principle, Pauli's exclusion principle and Hund's rule, electronic configuration of atoms, stability of half filled and completely filled orbitals.

**Unit III : Classification of Elements and Periodicity in Properties****08 Periods**

Significance of classification, brief history of the development of periodic table, modern periodic law and the present form of periodic trends in properties of elements : atomic radii, inert gas radii, ionization enthalpy, electron gain enthalpy, electronegativity, valency. Nomenclature of elements with atomic number greater than 100.

**Unit IV : Chemical Bonding and Molecular structure****14 Periods**

Valence electrons, ionic bond, covalent bond, bond parameters, Lewis structure, polar character of covalent bond, covalent character of ionic bond, valence bond theory, resonance, geometry covalent molecules, VSEPR theory, concept of hybridization, involving *s*, *p* and *d* orbitals and shape of some simple molecules, molecular orbital theory of homonuclear diatomic molecules (qualitative idea only), hydrogen bond.



**Unit V : States of Matter: Gases and Liquids****12 Period**

Three states of matter, intermolecular interactions, types of bonding, melting and boiling points, role of gas laws in elucidating the concept of the molecule, Boyle's law, Charles' law, Gay Lussacs law, Avogadro's law, ideal behaviour, empirical derivation of gas equation, Avogadro's number, ideal gas equation. Deviation from ideal behaviour, liquefaction of gases, critical temperature, kinetic energy and molecular speeds (elementary idea), Liquid State- vapour pressure, viscosity and surface tension (qualitative idea only, no mathematical derivations)

**Unit VI : Chemical Thermodynamics****14 Periods**

Concepts of System and types of systems, surroundings, work, heat, energy, extensive and intensive properties, state functions.

First law of thermodynamics : internal energy and enthalpy, heat capacity and specific heat, measurement of  $\Delta U$  and  $\Delta H$ , Hess's law of constant heat summation, enthalpy of bond dissociation, combustion, formation, atomization, sublimation, phase transition, ionization, solution and dilution. Second law of Thermodynamics (brief introduction)

Introduction of entropy as a state function, Gibb's energy change for spontaneous and non-spontaneous processes, criteria for equilibrium.

Third law of thermodynamics (brief introduction).

**Unit VII : Equilibrium****14 Periods**

Equilibrium in physical and chemical processes, dynamic nature of equilibrium, law of mass action equilibrium constant, factors affecting equilibrium : Le Chatelier's principle, ionic equilibrium- ionization of acids and bases, strong and weak electrolytes, degree of ionization, ionization of poly basic acids, acid strength, concept of pH, Henderson Equation, hydrolysis of salts (elementary idea), buffer solution, solubility product, common ion effect (with illustrative examples).

**Unit VIII: Redox Reactions****06 Periods**

Concept of oxidation and reduction, redox reactions, oxidation number, balancing redox reactions, in terms of loss and gain of electrons and change in oxidation number, applications of redox reactions.

**Unit IX : Hydrogen****08 Periods**

Position of hydrogen in periodic table, occurrence, isotopes, preparation, properties and uses of hydrogen, hydrides-ionic, covalent and interstitial; physical and chemical properties of water, heavy water, hydrogen peroxide-preparation, reactions and structure and use; hydrogen as a fuel.

**Unit X : *s*-Block Elements (Alkali and Alkaline Earth Metals) 10 Periods**

**Group 1 and Group 2 Elements :** General introduction, electronic configuration, occurrence, anomalous properties of the first element of each group, diagonal relationship, trends in the variation of properties (such as ionization enthalpy, atomic and ionic radii), trends in chemical reactivity with oxygen, water, hydrogen and halogens, uses.

**Preparation and Properties of Some Important Compounds :**

Sodium Carbonate, Sodium Chloride, Sodium Hydroxide and Sodium Hydrogencarbonate, Biological importance of Sodium and Potassium.

Calcium Oxide and Calcium Carbonate and their industrial uses, biological importance of Magnesium and Calcium.

**Unit XI : Some *p*-Block Elements****14 Periods****General Introduction to *p*-Block Elements :**

**Group 13 Elements :** General introduction, electronic configuration, occurrence, variation of properties, oxidation states, trends in chemical reactivity, anomalous properties of first element of the group, Boron-physical and chemical properties, some important compounds, Borax, Boric acid, Boron Hydrides, Aluminium: Reactions with acids and alkalies, uses.

**Group 14 Elements :** General introduction, electronic configuration, occurrence, variation of properties, oxidation states, trends in chemical reactivity, anomalous behaviour of first elements. Carbon-catenation, allotropic forms, physical and chemical properties; uses of some important compounds : oxides. Important compounds of Silicon and a few uses: Silicon Tetrachloride, Silicones, Silicates and Zeolites, their uses.



## **Unit XII : Organic Chemistry -Some Basic Principles and Technique**

**14 Periods**

General introduction, methods of purification, qualitative and quantitative analysis, classification and IUPAC nomenclature of organic compounds. Electronic displacements in a covalent bond: inductive effect, electromeric effect, resonance and hyper conjugation. Homolytic and heterolytic fission of a covalent bond: free radicals, carbocations, carbanions, electrophiles and nucleophiles, types of organic reactions.

## **Unit XIII: Hydrocarbons**

**12 Periods**

### **Classification of Hydrocarbons**

**Aliphatic Hydrocarbons :** Alkanes : Nomenclature, isomerism, conformation (ethane only), physical properties, chemical reactions including free radical mechanism of halogenation, combustion and pyrolysis.

Alkenes : Nomenclature, structure of double bond (ethene), geometrical isomerism, physical properties, methods of preparation, chemical reactions: addition of hydrogen, halogen, water, hydrogen halides (Markownikov's addition and peroxide effect), ozonolysis, oxidation, mechanism of electrophilic addition.

Alkynes : Nomenclature, structure of triple bond (ethyne), physical properties, methods of preparation, chemical reactions: acidic character of alkynes, addition reaction of - hydrogen, halogens, hydrogen halides and water.

Aromatic Hydrocarbons : Introduction, IUPAC nomenclature, benzene: resonance, aromaticity, chemical properties: mechanism of electrophilic substitution. Nitration, sulphonation, halogenation, Friedel Craft's alkylation and acylation, directive influence of functional group monosubstituted benzene. Carcinogenicity and toxicity.

## **Unit XIV : Environmental Chemistry**

**06Periods**

Environmental pollution : Air, water and soil pollution, chemical in atmosphere, smog, major atmospheric, acid rain, ozone and its reactions, effects of depletion of ozone layer, greenhouse effect and global warming : pollution due to industrial wastes, green chemistry as an alternative tool for reducing pollution, strategies for control of environmental pollution.

## Chemistry (Code No. 043)

### Question Paper Design

Class : XI (2016-17)

Time : 3 Hours]

[Max. Marks : 70

S. Typology of Questions	V.S.A.	S.A.-I	S.A.-II	V.B.Q.	L.A.	Total Marks	% Weightage
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- |    |  |  |  |  |  |  |  |
|----|--|--|--|--|--|--|--|
| 1. | Remembering-(Knowledge based Since recall questions, to know specific facts, terms, concepts, principles, or theories, Identify, define, or recite, information)   |  |  |  |  |  |  |
| 2. | Understanding Comprehension-to be familiar with meaning and to understand conceptually, interpret, compare, contrast, explain, paraphrase information.   |  |  |  |  |  |  |
| 3. | Application (Use abstract information in concrete situation, to apply knowledge to new situations, Use given content to interpret a situation, provide an example, or solve a problem)   |  |  |  |  |  |  |
| 4. | High order Thinking Skills (Analysis and Synthesis : Classify, compare, contrast, or differentiate between different pieces of information, organize and/or integrate unique pieces of information from a variety of sources.) |  |  |  |  |  |  |
| 5. | Evaluation A: (Appraise, Judge, and/or justify the value or worth of a decision or outcome, to to predict outcomes based on values)  |  |  |  |  |  |  |

### Question Wise Break Up

Type of Ques.	Mark per Ques.	Total No. of Ques.	Total Marks
VSA	1	5	05
SA-I	2	5	10
SA-II	3	12	36
VBQ	4	1	04
LA	5	3	15
<b>Total</b>		26	70

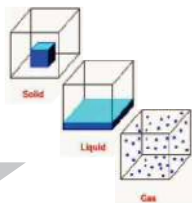
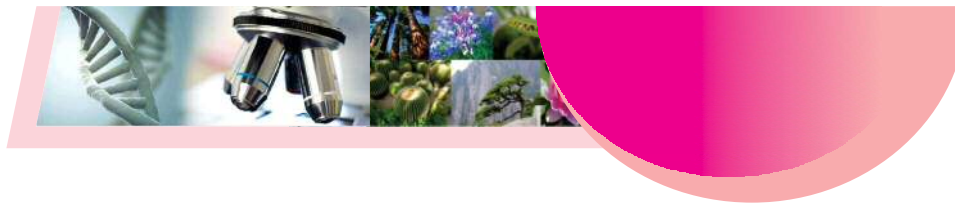
- 1. Internal Choice :** There is no overall choice in the paper. However, there is an internal choice in one question of 2 marks weightage, one question of 3 marks weightage and all the three questions of 5 marks weightage.
- The above template is only a sample. Suitable internal variations may be made for generating similar templates keeping the overall weightage to different form of questions and typology of questions same.

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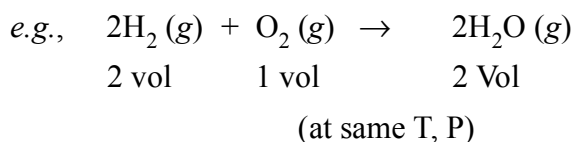
## Chapter - 1

# Some Basic Concepts Of Chemistry

- **Matter** : Anything that has mass and occupy space.
- **Precision** : It refers to the closeness of various measurements for the same quantity.
- **Accuracy** : It refers to the agreement of a particular value to the true value of the result.
- **Mass and weight** : Mass of a substance is the amount of matter present in it while weight is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.
- **Volume** :  $1 \text{ L} = 1 \text{ dm}^3 = 10^3 \text{ cm}^3 = 10^{-3} \text{ m}^3$
- **Temperature** :  $K = ^\circ\text{C} + 273.15$ ;  $\frac{^\circ\text{F} - 32}{9} = \frac{^\circ\text{C}}{5}$
- **Standard Temperature Pressure (STP)** :  $0^\circ\text{C}$  (273.15 K) temperature and 1 atm pressure.
- **Normal Temperature Pressure (NTP)** :  $20^\circ\text{C}$  (293.15 K) temperature and 1 atm pressure.
- **Standard Ambient Temperature Pressure (SATP)** :  $25^\circ\text{C}$  (298.15 K) temperature and 1 atm pressure
- **Scientific Notation** : Expressing a number in the form  $N \times 10^n$ , and N can vary b/w 1 to 10.
- **Significant figures** : These are meaningful digits which are known with certainty.
- **Laws of Chemical Combination** :
  - **Law of Conservation of Mass** (Antonie Lavoisier) : Mass can neither be created nor be destroyed.
  - **Law of Definite Proportions** (Joseph Proust) : A given compound

always contains the same elements in the same proportion by mass.

- **Law of Multiple Proportions** (John Dalton) : When two elements combine to form two or more compounds, then the different masses of one element, which combine with a fixed mass of the other, bear a simple ratio to one another.
- **Gay Lussac's Law** : When gases combine or are produced in a chemical reaction, they do so in a simple ratio provided all gases are in the same temperature and pressure.



- **Atomic Mass** : It is defined as the average relative mass of an atom of an element as compared to the mass of an atom of carbon – 12 taken as 12. Atomic mass is represented by 'u' (unified mass).

$$1u = 1.66056 \times 10^{-24} \text{ g}$$

- **Molecular mass** : It is the sum of the atomic mass of the elements present in the molecule.

*For example* : Molecular mass of  $\text{CH}_4 = (1 \times 12) + (4 \times 1) = 16 \text{ u}$

- **Avogadro Number** : It is the amount of atoms or molecules present in one mole of a substance.

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

- **Molar Mass** : The mass of one mole of a substance in grams is called its molar mass.

*For example* : Molar mass of  $\text{CH}_4 = (1 \times 12) + (4 \times 1) = 16 \text{ g mol}^{-1}$

- **Mole (n)** : It is amount of a substance that contains as many particles or entities as the number of atoms in exactly 12 grams of pure C-12.

1 mole of a substance = Molar mass of substance = Avogadro's Number of chemical units = 22.4L volume at STP of gaseous substance

*e.g.*, 1 mole of  $\text{CH}_4 = 16 \text{ g}$  of  $\text{CH}_4 = 6.022 \times 10^{23}$  molecules of  $\text{CH}_4 = 22.4 \text{ L}$  at STP

$$n = \frac{wg}{M_m} = \frac{VL \text{ (at STP)}}{22.4L} = \frac{x \text{ particles}}{N_A} = \frac{MV}{1000}$$

- **Molar Volume ( $V_m$ )** : It is volume occupied by one mole of any substance. Molar volume of a gas = 22.4L at STP (273 K, 1atm) or 22.7L at STP (273

K, 1 bar)

Calculating Molar Volume:  $PV = nRT$

$$\therefore V = \frac{nRT}{P} = \frac{1 \text{ mol} \times 0.082 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}}{1 \text{ atm}} = 22.4 \text{ L}$$

Or

$$V = \frac{nRT}{P} = \frac{1 \text{ mol} \times 0.083 \text{ L bar K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}}{1 \text{ atm}} = 22.7 \text{ L}$$

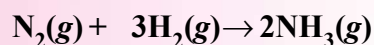
- **Percentage Composition** : Mass % of the element

$$= \frac{\text{Mass of element in a molecule of the compound} \times 100}{\text{Molecular mass of the compound}}$$

- **Empirical Formula** : It represents the simplest whole number ratio of various atoms present in a compound. *e.g.*, CH is the empirical formula of benzene.
- **Molecular Formula** : It shows the exact number of different of atoms present in a molecule of a compound. *e.g.*, C<sub>6</sub>H<sub>6</sub> is the molecular formula of benzene.
- **Relationship between empirical and molecular formulae** : Molecular formula =  $n \times$  Empirical formula

Where; 
$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

- **Information Conveyed by a chemical equation** :



- (i) 1 molecule of N<sub>2</sub> + 3 molecules of H<sub>2</sub> → 2 molecules of NH<sub>3</sub>
- (ii) 1 mole of N<sub>2</sub> + 3 mole of H<sub>2</sub> → 2 mole of NH<sub>3</sub>
- (iii) 1 × 28g of N<sub>2</sub> + 3 × 2 g of H<sub>2</sub> → 2 × 17 g of NH<sub>3</sub>
- (iv) 1 × 22.4L of N<sub>2</sub> + 3 × 22.4L of H<sub>2</sub> → 2 × 22.4L of NH<sub>3</sub>  
at STP at STP at STP

- **Limiting Reagent** : It is the reactant which gets consumed first or limits the amount of product formed.
- **Mass Percent** : It is the mass of the solute in grams per 100 grams of the solution.

$$\text{Mass percent} = \frac{\text{Mass of solute in g} \times 100}{\text{Mass of solution in g}}$$

- **Parts per million (ppm)** : It is part of solute per million part of solution by mass.

$$\text{ppm} = \frac{\text{Parts of solute (by mass)} \times 10^6}{\text{Parts of solution (by mass)}}$$

- **Molarity (M)** : It is number of moles of solute dissolved per litre (dm<sup>3</sup>) of the solution.

$$\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in L}}$$

$$\text{Molarity equation : } M_1 V_1 = M_2 V_2$$

(Before dilution) (After Dilution)

Molarity of a solution decreases on increasing temperature.

Molarity of pure water is 55.56 mol L<sup>-1</sup>

- **Molality (m)**—It is number of moles of solute dissolved per 1000g (1kg) of solvent.

$$\text{Molality} = \frac{\text{No. of moles of solute}}{\text{mass of solution in kg}}$$

Molality is independent of temperature.

- **Mole Fraction(x)** is the ratio of number of moles of one component to the total number of moles (solute and solvents) present in the solution.

$$x_1 = \frac{n_1}{n_1 + n_2} \quad \text{and} \quad x_2 = \frac{n_2}{n_1 + n_2}$$

The sum of all the mole fractions in a solution is equal to one. *i.e.*,  $x_1 + x_2 = 1$

## Importance of Chemistry & Nature of Matter

### 1-Mark Questions

1. Name two chemical compounds used in treatment of cancer.
2. What is AZT ? Write its use.
3. Give an example each of homogeneous and heterogeneous mixture.
4. Differentiate solids, liquids & gases in terms of volume & shapes.
5. Classify following as pure substances and mixtures : Air, glucose, gold,

sodium and milk.

6. What is the difference between molecules and compounds? Give examples of each.

### Properties of matter and their Measurement

7. What is the SI unit of density ?  
8. What is the SI unit of molarity ?  
9. Define accuracy and precision.  
10. What are the two different system of measurement ?  
11. What is the difference between mass & weight ?

### Uncertainty in Measurement

12. Define significant figures.  
13. Define accuracy and precision  
14. Which measurement is more precise 4.0g or 4.00g ? [Ans. 4.00 g]  
15. How many significant figures are there in (i) 3.070 and (ii) 0.0025 ?  
[Ans. (i) 4 (ii) 2]  
16. Express the following in the scientific notation : (i) 0.0048 (ii) 234,000

### Laws of Chemical Combinations & Dalton's Atomic Theory

17. State Avogadro's law.  
18. State law of definite proportions.  
19. State Gay Lussac's Law of combining volumes of gases.  
20. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how much volume of water vapour would be produced ?  
[Ans. 10 volumes]

### Atomic and Molecular masses and Mole Concept

21. Define unified mass (u).  
22. Calculate the number of atoms in 32.0 u of He. [Ans. 8]  
23. Define molar volume of a gas.  
24. What is the volume of 17 g of  $\text{NH}_3$  gas at STP (298 K, 1 atm) ?  
[Ans. 22.4 L]  
25. What is the value of one mole ?

26. Calculate the number of molecules present in 22.0 g of  $\text{CO}_2$ .  
[Ans.  $3.011 \times 10^{23}$ ]
27. How many molecules of  $\text{SO}_2$  are present in 11.2 L at STP ?  
[Ans.  $3.011 \times 10^{23}$ ]
28. Which has more number of atoms ? 1.0 g Na or 1.0 g Mg. [Ans. 1.0 g Na]
29. How many oxygen atoms are present in 16 g of ozone ( $\text{O}_3$ ) ?  
[Ans.  $2.007 \times 10^{23}$ ]
30. At STP, what will be the volume of  $6.022 \times 10^{23}$  molecules of  $\text{H}_2$  ?  
[Ans. 22.4L]
31. 1L of a gas at STP weighs 1.97g. What is molecular mass ?  
[Ans. 44.128 g  $\text{mol}^{-1}$ ]

#### Percentage Composition, Empirical and Molecular Formula

32. Write the relationship between empirical formula and molecular formula.
33. Which is more informative ? Empirical formula or Molecular formula.
34. A substance has molecular formula  $\text{C}_6\text{H}_{12}\text{O}_6$ . What is its empirical formula.
35. Empirical formula of a compound X (Molar mass = 78  $\text{mol}^{-1}$ )  $\text{CH}$ . Write its molecular formula.

#### Stoichiometry and Stoichiometric Calculations

36. How are 0.5 mol  $\text{Na}_2\text{CO}_3$  and 0.5 M  $\text{Na}_2\text{CO}_3$  different from each other ?
37. Why molality is preferred over molarity of a solution ?
38. Define molarity of a solution.
39. What is the effect of temperature on molarity of solution ?
40. What is limiting reactant in a reaction ?

### Importance of Chemistry & Nature of Matter

#### 2 Mark Questions

1. How can we say that sugar is solid and water is liquid?
2. How is matter classified at macroscopic level?
3. Classify following substances as element, compounds and mixtures : water, tea, silver, steel, carbon dioxide and platinum.

### Properties of matter and their Measurement

4. The body temperature of a normal healthy person is  $37^{\circ}\text{C}$ . Calculate its value in  $^{\circ}\text{F}$ .
5. At what temperature will both the Celsius and Fahrenheit scales read the same value?
6. Convert 5L into  $\text{m}^3$ .
7. What do the following prefixes stand for : (a) pico (b) nano (c) micro (d) deci

### Uncertainty in Measurement

8. How many significant figures are present in the answer of the following calculations :  
(i)  $0.0125 + 0.8250 + 0.025$                       (ii)  $\frac{0.025 \times 298.15 \times .1155}{0.5785}$
9. Convert '450 pm' into SI unit and write the answer in scientific notation upto 2 significant figures.

[Ans.  $4.5 \times 10^{-10} \text{ m}$ ]

10. The density of vanadium is  $5.96 \text{ g cm}^{-3}$ . Express this in SI unit.

[Ans.  $5960 \text{ kg m}^{-3}$ ]

### Laws of Chemical Combinations & Dalton's Atomic Theory

11. 45.4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed. The reaction is given below :  $2 \text{ N}_2 (\text{g}) + \text{O}_2 (\text{g}) \rightarrow 2 \text{ N}_2\text{O} (\text{g})$  Which law is being obeyed in this experiment? Write the statement of the law.
12. Write main points of Dalton's Atomic Theory.

### Atomic and Molecular masses and Mole Concept

13. Give one example each of a molecule in which empirical formula and molecular formula is (i) Same (ii) Different.
14. Calculate the number of moles in the following masses :  
(i) 7.85g of Fe; (ii) 7.9mg of Ca
15. Calculate average atomic mass of chlorine using following data:

Isotope	% Natural abundance	Molar mass
$^{35}\text{Cl}$	75.77	34.9689
$^{37}\text{Cl}$	24.33	36.9659 [Ans. 35.5 u]

### Percentage composition, empirical and molecular formula

16. Give one example of molecule in which empirical formula and molecular formula are (i) same (ii) different.
17. Calculate the present of carbon, hydrogen and oxygen in ethanol ( $\text{C}_2\text{H}_5\text{OH}$ )  
[Ans. 52.14%, 13.13%, 34.73%]
18. How much copper can be obtained from 100 g of  $\text{CuSO}_4$  ? [Ans. 39.8g]

### Stoichiometry and Stoichiometric Calculations

19. Calculate the amount of water (g) produced by the combustion of 16 g of methane.  
[Ans. 36g]
20. How many moles of methane are required to produce 22 g  $\text{CO}_2$  (g) after combustion?  
[Ans. 0.5 mol]
21. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute.  
[Ans. 10%]
22. Calculate molarity of water if its density is  $1.00 \text{ g mL}^{-1}$ . [Ans. 55.56 M]
23. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution. [Ans. 0.4 M]
24. The density of 3 M solution of NaCl is  $1.25 \text{ g mL}^{-1}$ . Calculate molality of the solution.  
[Ans. 2.8m]
25. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).  
[Ans. 2.31 M]
26.  $\text{NH}_3$  gas can be prepared by Haber's process as,  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ . At a particular moment concentration of all the species is 2 moles; calculate the concentration of  $\text{N}_2$  and  $\text{H}_2$  taken initially.  
[Ans. 3 mole, 5 moles]
27. A sample of drinking water was found to be severely contaminated with



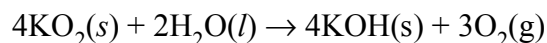
chloroform,  $\text{CHCl}_3$ , supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in percent by mass.

(ii) Determine the molality of chloroform in the water sample.

[Ans. (i)  $\sim 15 \times 10^{-14}$  g (ii)  $1.25 \times 10^{-4}$  m]

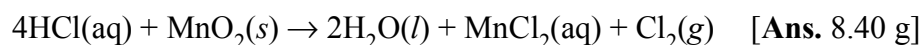
28. Potassium superoxide,  $\text{KO}_2$  is used in rebreathing gas masks to generate oxygen.



If a reaction vessel contains 0.15 mol  $\text{KO}_2$  and 0.10 mol  $\text{H}_2\text{O}$ , what is the limiting reactant? How many moles of oxygen can be produced?

[Ans.  $\text{KO}_2$ , 1.125 mol]

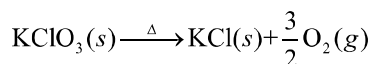
29. How many grams of  $\text{HCl}$  react with 5.0 g of  $\text{MnO}_2$  according to the equation.



30. 0.5 mol of  $\text{H}_2\text{S}$  and  $\text{SO}_2$  are mixed together in a reaction flask in which the following reaction takes place :  $2\text{H}_2\text{S}(g) + \text{SO}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 3\text{S}(s)$

Calculate the number of moles of sulphur formed. [Ans. 0.75 mol]

31. Pure oxygen is prepared by thermal decomposition of  $\text{KClO}_3$  according to the equation :



Calculate the volume of oxygen gas liberated at STP by heating 12.25 g  $\text{KClO}_3(s)$ . [Ans. 3.36 L]

## Importance of Chemistry and Nature of Matter

### 3-Marks Questions

1. Give three main points of difference between a compound and a mixture.
2. Define homogeneous and heterogeneous mixture with example.

### Properties of matter and their Measurement

3. Write seven fundamental quantities & their units
4. Pressure is defined as force per unit area of the surface. The SI unit of

pressure, Pascal is :

$$1\text{Pa} = 1\text{Nm}^{-2}$$

If mass of air at sea level is  $1034\text{ g cm}^{-2}$ , calculate the pressure in Pascal.

[Ans.  $1.01332 \times 10^5\text{ Pa}$ ]

### Laws of Chemical Combinations & Dalton's Atomic Theory

5. The following data are obtained when dinitrogen and dioxygen react together to form different compounds :

	(i)	(ii)	(iii)	(iv)
Mass of dinitrogen	14	14	28	28
Mass of dioxygen	16	32	32	80

Which law of chemical combination is obeyed by the above experimental data ? Give its statement.

### Atomic and Molecular Masses and Mole Concept

6. Calculate :

- (i) Mass in gram of  $5.8\text{ mol N}_2\text{O}$
- (ii) No of moles in  $8.0\text{ g of O}_2$
- (iii) Molar mass if  $11.2\text{ L at STP weigh } 8.5\text{ g}$ .

[Ans. (i)  $255.2\text{ g}$  (ii)  $0.25\text{ mol}$  (iii)  $17\text{ g mol}^{-1}$ ]

7. In three moles of ethane ( $\text{C}_2\text{H}_6$ ), calculate the following :

- (i) No of moles of carbon atom,
- (ii) No of moles of hydrogen atoms,
- (iii) No of molecules of ethane.

[Ans. (i)  $6\text{ moles}$ , (ii)  $18\text{ moles}$ , (iii)  $1.81 \times 10^{24}$ ]

8.  $16\text{ g}$  of an ideal gas  $\text{SO}_x$  occupies  $5.6\text{ L}$  at STP. What is its molecular mass? What is the value of  $X$  ?

[Ans.  $64u$ ,  $x = 2$ ]

9. Calculate the number of moles :

- (i)  $5.0\text{ L of } 0.75\text{ M Na}_2\text{CO}_3$
- (ii)  $7.85\text{ g of Fe}$
- (iii)  $34.2\text{ of sucrose (C}_{12}\text{H}_{22}\text{O}_{11})$

[Ans. (i)  $3.75$ , (ii)  $0.14$ , (iii)  $0.1$ ]

10. Calculate the number of atoms in each of the following :

- (i)  $52\text{ moles of Ar}$ . (ii)  $52u\text{ of He}$  (iii)  $52\text{g of He}$ .

[Ans. (i)  $3.13 \times 10^{25}$  (ii)  $13$  (iii)  $7.83 \times 10^{24}$ ]

### Percentage Composition, Empirical and Molecular Formula

11. Vitamin C is essential for the prevention of scurvy. Combustion of 0.2000g of vitamin C gives 0.2998g of  $\text{CO}_2$  and 0.819g of  $\text{H}_2\text{O}$ . What is the empirical formula of vitamin C ?
12. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas? [Ans.  $\text{CH}_2\text{Cl}$ ,  $\text{C}_2\text{H}_4\text{Cl}_2$ ]
13. A compound made up of two elements A and B has A = 70%, B = 30%. Their relative number of moles in the compound is 1.25 and 1.88, calculate :  
(i) Atomic masses of the elements A and B  
(ii) Molecular formula of the compound , if its molecular mass is found to be 160. [Ans. (i) 56 and 16, (ii)  $\text{A}_2\text{B}_3$ ]

### Stoichiometry and Stoichiometric Calculations

14. Calculate the mass of sodium acetate ( $\text{CH}_3\text{COONa}$ ) required making 500 mL of 0.375 molar aqueous solution. (Molar mass of sodium acetate is  $82.0245 \text{ g mol}^{-1}$ ). [Ans. 15.375 g]
15. Calculate the concentration of nitric acid in moles per litre in a sample which has a density,  $1.41 \text{ g mL}^{-1}$  and the mass per cent of nitric acid in it being 69%. [Ans. 15.44 M]
16. What is the concentration of sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ) in  $\text{mol L}^{-1}$  if its 20 g are dissolved in enough water to make a final volume up to 2L ? [Ans. 0.029 M]
17. Calcium carbonate reacts with aqueous  $\text{HCl}$  according to the reaction :  
 $\text{CaCO}_3 (\text{s}) + 2 \text{HCl} (\text{aq}) \rightarrow \text{CaCl}_2 (\text{aq}) + \text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{l})$   
What mass of  $\text{CaCO}_3$  is required to react completely with 25 mL of 0.75 M  $\text{HCl}$  ? [Ans. 0.94 g]
18. The reaction  $2\text{C} + \text{O}_2 \rightarrow 2\text{CO}$  is carried out by taking 24.0 g of carbon and 96.0 g of  $\text{O}_2$ . Find out.  
(i) Which reactant is left in excess ?  
(ii) How much of it is left ?  
(iii) How many grams of the other reactant should be taken so that nothing is left at the end of the reaction ? [Ans. (i)  $\text{O}_2$ , (ii) 64 g, (iii) 72]

### HOTS Question

19. A 10 g sample of a mixture of calcium chloride and sodium chloride is treated with  $\text{Na}_2\text{CO}_3$  to precipitate calcium as calcium carbonate. This  $\text{CaCO}_3$  is heated to convert all the calcium to  $\text{CaO}$  and the final mass of

CaO is 1.62 g. Calculate % by mass of NaCl in original solution.

[Ans. 67.9%]

## Atomic and Molecular masses and Mole Concept

### 5-Mark Questions

1. (i) A black dot used as a full stop at the end of a sentence has a mass of about one attogram. Assuming that the dot is made up of carbon, calculate the approximate number of carbon atoms present in the dot.

[Hint : 1 attogram =  $10^{-18}$  g] [Ans.  $5.02 \times 10^4$ ]

- (ii) Which one of the following will have largest number of atoms ?

(a) 1g Au (s) (b) 1g Na (s) (c) 1g Li (s) (d) 1g of  $\text{Cl}_2$ (g)

[Ans. (i) 39.81 g (ii) 1 g of Li]

### Percentage Composition, Empirical and Molecular Formula

2. (i) What is the difference between empirical formula and molecular formula ?

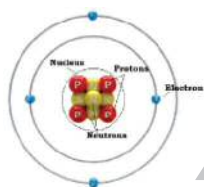
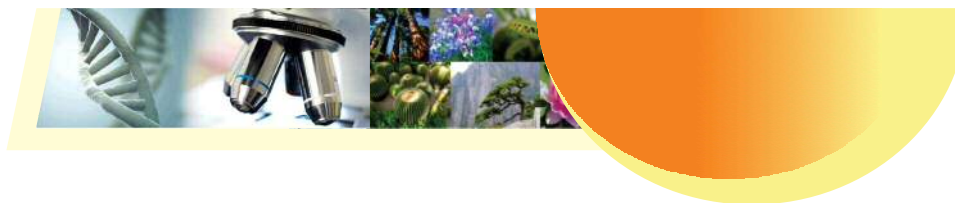
(ii) A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

[Ans. (i)  $\text{CH}$ , (ii)  $26 \text{ g mol}^{-1}$ , (iii)  $\text{C}_2\text{H}_2$ ]

### Stoichiometry and Stoichiometric Calculations

3. (i) What is the difference between Molarity and Molality.  
(ii) The Molarity of a solution of sulphuric acid is 1.35 M. Calculate its molality. (The density of acid solution is  $1.02 \text{ g cm}^{-3}$ ). [Ans. 1.52 m]

4. (i) Define : (a) Mole fraction (b) Mass percentage.  
(ii) If the density of methanol is  $0.793 \text{ kg L}^{-1}$ , what is its volume needed for making 2.5 L of its 0.25 M solution ? [Ans. 0.0025 L]



## Chapter - 2

# Atomic Structure

**(1) Atom (Given by Dalton) :** Matter is made up of extremely small particles which are indivisible in nature. It consists of subatomic particles electron, proton and neutrons known as **FUNDAMENTAL PARTICLES**.

**(2) Electron (Named by Stoney) :** Discovered by Cathode Ray experiment [In CROOK'S TUBES]. A long glass tube with two metal electrodes. At every low pressure when high voltage is applied a flow is produced due to flow of – ve charge particle [KNOWN AS ELECTRON], cathode rays. Cathode rays have – ve charge, travel in straight lines have electric and magnetic field have heating effect more penetrating effect.

Charge on  $e^-$  was found by **OIL DROP** experiment [MILLIKAN].

**(3) Proton (Discovered by GOLDSTEIN in ANODE RAY EXPT) :** in a perforated cathode tube with gas at low pressure high voltage was passed b/w electrode rays from cathode produced green **FLOURESCENCE** on ZnS wall. These were called as **ANODE RAYS**. They travel in straight line, with + ve charge, get deflected in electric and magnetic field.

**(4) Neutron :** Fundamental particle which carries **no** charge but has mass equal to N atom or **PROTON**. Discovered by **JAMES CHADWICK**.

Name of Constant	UNIT	Electron	Proton	Neutron
Mass	amu	0.000546	1.00728	1.008665
	Kg	$9.109 \times 10^{-31}$	$1.673 \times 10^{-27}$	$1.675 \times 10^{-27}$
Charge	Coloumbs	$1.602 \times 10^{-19}$	$+ 1.602 \times 10^{-19}$	Zero
	esu	$- 4.8 \times 10^{-10}$	$+ 4.8 \times 10^{-10}$	Zero
	Relative	- 1	+ 1	Zero

**(5) J.J. Thomson :** The sphere of + ve charge nucleus model of atom is

balanced by coulombic force of attraction of  $e^-$ . Like a Raisin Pudding Model.

(6) **Rutherford** (Discovery of nucleus) :  ${}_2^4\text{He}$   $\alpha$  particles (+ ve charge) bombarded on gold foil.

- (a) 99.9% passed without deflection : Most space inside the atom is empty.
- (b) Only few deflected therefore mass of atom centrally placed called nucleus
- (c) very few deflected back therefore mass of atom contains + ve charge particles [Protons].
- (d) Atom is electrically neutral hence – ve charge particles placed outside the nucleus and have very less mass.

**Limitations** : No distribution and energies of  $e^-$  considered, could not explain  $e^-$  does not fall into the nucleus or not; no details of line spectra of H atom.

(7) (a) Atomic number (Z) : The no. of protons or electron in a neutral atom  
or

No. of protons in an atom (or ion).

(b) At mass no. (A) : Total no. of protons and neutron in an atom

$${}_Z^AX \quad A = \text{No. of } (P + n) \quad A - Z = \text{No. of neutrons}$$

(8) (a) Isotopes : Atoms of same element with different mass no.  $[_7^{14}\text{N}, _7^{15}\text{N}]$

(b) Isobars : Atoms of different element with same mass no.  $[_{20}^{40}\text{Ca}, _{18}^{40}\text{Ar}]$

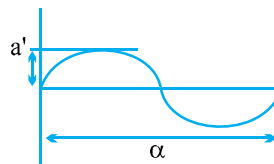
(c) Isotones : Atoms of different element with same no. of neutron  $[_{11}^{23}\text{Na}, _{12}^{24}\text{Mg}]$

(d) Isoelectronic : Atoms, molecules or ions with same no. of  $e^-$   $[\text{Ne}; \text{O}^{2-}]$ .

(9) **Electromagnetic radiations** : Energy emitted from any source (in forms of waves) in which electric and magnetic fields oscillated perpendicular to each other and travelling with a velocity to light is known as EM radiation.

(10) **Characteristics of waves** :

- (a) Wavelength : the distance of one crest and one trough in a wave. Denoted by ' $\lambda$ '.
- (b) Frequency : no. of waves passing through a given point in one second.



Denoted by  $\nu$ . 
$$\left[ \begin{array}{l} \nu = \frac{1}{t} \Rightarrow \text{sec}^{-1} \text{ or Hz} \\ t = \text{Time period} \end{array} \right]$$

(c) Amplitude : The height of crest or depth of a trough denoted by 'a'.

(d) Wave no. : No. of waves per unit length denoted by  $\bar{\nu}$

$$\bar{\nu} = \frac{1}{\lambda} = \text{cm}^{-1} \text{ (or m}^{-1}\text{)}$$

(e) Velocity : Linear distance travelled by a wave in one second.

$$\begin{aligned} \text{velocity of light } c &= \frac{\text{Distance}}{\text{Time}} = \lambda \times \frac{1}{t} = \nu \lambda \\ \therefore \nu &= \frac{c}{\lambda} \end{aligned}$$

**(10) Energywise order for EM radiation.**

cosmic <  $\gamma$  rays < X rays < uv < VIBGYOR < IR < Microwaves < Radiowaves

$$\begin{array}{ccccccc} \longrightarrow & \longrightarrow & \longrightarrow & \longrightarrow & \longrightarrow & & \\ \lambda \text{ Inc.} & \nu \text{ Dec.} & & \bar{\nu} \text{ Dec.} & \text{Energy Dec.} & & \end{array}$$

**(11) Proton** : A packet or particle of light energy is known as **Photon**.

**(12) Planck's quantum theory** : The energy emitted from a source or absorbed by a source is not continuous but discontinuous in form of small packet of energy [for light each packet is known as photon]

Energy [emitted or absorbed] by a photon  $\propto \nu$  (frequency)

$E \propto \nu$  or  $E = h\nu$  [ $h$  = Planck's constant]

$$= 6.626 \times 10^{-34} \text{ J sec}$$

If 'n' photons are emitted  $E = nh\nu$

**(13) Photo electric effect** : The phenomenon of ejection of electrons from a metal surface when a light radiation of suitable frequency falls on metal surface.

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

$h\nu \Rightarrow$  Light radiation falling upon metal surface.

$h\nu_0 \Rightarrow$  Energy used for work function or [Energy for removing  $e^-$  from metal]

$$\frac{1}{2}mv^2 = K. \text{ energy by which } e^- \text{ is emitted from metal surface.}$$

(a) Large atoms have less work function.

$\therefore e^-$  emitted with more velocity.

(b) Small atoms have more work function.

$\therefore e^-$  emitted with less velocity.

**Debroglie equation :** All material particles possess both matter should also exhibit wave like properties. Wave character as well as wave character as well as wave character.

$$\lambda \propto \frac{1}{\text{Momentum}}$$
$$\lambda = \frac{h}{mv}$$

For microscopic particles mass is very less therefore ' $\lambda$ ' more and more wave character.

For macroscopic particles mass is large  $\lambda$  is less therefore more particle character.

Dual behaviour  $\left[ \begin{array}{l} e^- \text{ behaves as a particle as well as wave} \\ \text{Photon behaves as wave as well as particle} \end{array} \right]$

### Heisenberg's Uncertainty Principle

It is impossible simultaneously to determine the exact position and exact velocity of a subatomic particle.

$$\Delta x \times m\Delta v \geq \frac{h}{4\pi}$$

$\Delta x$  = uncertainty in position  
 $\Delta v$  = uncertainty in momentum

For microscopic (mass very less) certainty in position is less therefore  $\Delta x$  is more  $\Delta v$  is less.

For macroscopic (large mass) certainty in position is more  $\Delta x$  is less  $\Delta v$  is more.

**Bohr's theory** for H [H like one  $e^-$  systems  $\text{He}^+$ ;  $\text{Li}^{2+}$ ]

(1)  $e^-$  revolving round the nucleus in circular path [stationary state; SHELL]

with a definite angular momentum  $\frac{nh}{2\pi}$  [n no. of shell of  $e^-$ ] and with definite

energy  $\left[ \frac{-2\pi^2 me^4 Z^2}{n^2 h^2} \right] \Rightarrow -2.18 \times 10^{-18} \frac{Z^2}{n^2} \text{ J/Atom.}$



(2) As  $n$  increases,  $Z$  Decreases Energy of  $e^-$  becomes  $\leq$  less – ve [Due to less force of Proton attraction]

As  $n$  Decreases  $Z$  increases Energy of  $e^-$  becomes More – ve [Due to more force of attraction by protons]

(3) In infinity shell  $e^-$  has zero force of attraction therefore zero energy.

(4) Electron energy only changes by definite values  $\Delta E = E_f - E_i$ .

$$\frac{hc}{\lambda} = 2.18 \times 10^{-18} \left[ \frac{1}{n_i^2} - \frac{1}{n_f^2} \right] Z^2 \text{ J/Atom}$$

$$\frac{1}{\lambda} = \frac{2.18 \times 10^{-18}}{hc} \left[ \frac{1}{n_i^2} - \frac{1}{n_f^2} \right] Z^2$$

$$\left[ \begin{array}{l} \text{If } n_i > n_f \text{ energy emitted by } e^- \\ n_i < n_f \text{ energy absorbed by } e^- \end{array} \right]$$

$$= 109678 \left[ \frac{1}{n_i^2} - \frac{1}{n_f^2} \right] Z^2 \text{ cm}^{-1}$$

Rydberg constant

**Hydrogen spectrum :** When  $e^-$  in hydrogen atom is provided energy it gets excited to higher shell from ground state, it comes back to ground state by emitting energy in definite values.

**“Quanta” :** The emission of light energy is known as emission spectra. It corresponds to each atom depending upon which energy shell  $e^-$  is excited. It is **discontinuous** spectra as ‘ $\lambda$ ’ of light radiations do not merge with each other like is VIBGYOR (Continuous Spectra).

When  $e^-$  falls from any excited state to

$$\frac{1}{\lambda} = 1,09,678 \left[ \frac{1}{n_f^2} - \frac{1}{n_i^2} \right] Z^2$$

- (a) 1<sup>st</sup> energy level  $n_f = 1$ ,  $n_i = 2, 3, 4, \dots$  [Lyman series] (UV)
- (b) When  $e^-$  to final state  $n = 2$ ,  $n_i = 3, 4, 5, \dots$  [Balmer series] (VIBGYOR)
- (c) When  $e^-$  to falls to final state 3  $n_i = 4, 5, 6$  [Paschem series] IR.
- (d) When  $e^-$  to falls to final state 4  $n_i = 5, 6, 7$  [Bracket series] IR.
- (e) When  $e^-$  to falls to final state 5  $n_i = 6, 7, 8$  [Pfund series] IR.

**Quantum numbers :** The no. which **completely** define the **state** of  $e^-$ .

(a) Position and energy of  $e^-$  (b) Path of  $e^-$  (c) Orientation in space (d) Rotational motion of  $e^-$ . These are described by four quantum numbers.





**(1) Principal Q. No. :** It describes the distance of  $e^-$  from **nucleus ‘ $n$ ’** i.e., defines the **shell** no. It is denoted by ‘ $n$ ’.

$$n = 1, 2, 3, 4, 5, \dots$$

K, L, M, N, O .....

**(2) Azimuthal Q. No. :** It defines the path of  $e^-$  decided by angular momentum of  $e^-$ . Each angular momentum value corresponds to one subshell. The no. of subshells in a shell is 0 to  $n - 1$ .

$$n \quad l \text{ (0 to } n-1 \text{)}$$

1	0	$l = 0$	‘s’	subshell	
2	0, 1	$l = 1$	‘p’	subshell	
3	0, 1, 2	$l = 2$	‘d’	subshell	
4	0, 1, 2, 3	$l = 3$	‘f’	subshell	

All subshells are wave functions for locating  $e^-$ .

In the same shell energy wise  $S < P < d < f$ .

**(3) Magnetic Q. No. :** It gives the no. of magnetic orientations an  $e^-$  can have in a subshell. The no. of magnetic orientation an  $e^-$  can have in a subshell  $\Rightarrow -l$  to 0 to  $+l$ .

**(4) Spin Q. No. :** An  $e^-$  is continuously spinning on its own axis. This Q. No.

describes  $e^-$  can have clockwise spin motion  $\left(+\frac{1}{2} \text{ value}\right)$  or  $e^-$  can have anticlockwise spin motion  $\left(-\frac{1}{2}\right)$ . An orbital can have maximum two  $e^-$

one with clockwise and other with anticlockwise spin.

### Aufbau principle

- $e^-$  are filled in increasing order of energy of subshell.
- As ‘ $n + l$ ’ value inc. energy of  $e^-$  increases in that subshell.
- For two subshells with same ‘ $n + l$ ’ value. As ‘ $n$ ’ value increases energy of  $e^-$  increases.

### Pauli’s principle

No two  $e^-$  can have same set of 4 quantum nos. If two  $e^-$  are present in

same shell, subshell, orbital they will have different spin value.

### Hund's rule

The pairing of  $e^-$  in degenerate orbitals (different orbitals with same energy) will get paired only once they have been singly occupied.

### Important Points

The filling of  $e^-$  in subshells follows this order. (As per Aufbau principle)

(A)  $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p$

(B) Half filled and completely filled subshells have more **stability** than incompletely filled subshells.

(24)  $\text{Cr} = [\text{Ar}] 4s^2 3d^4$  changes to  $\text{Cr} = [\text{Ar}] 4s^1 3d^5$

(29)  $\text{Cu} = [\text{Ar}] 4s^2 3d^9$  changes to  $\text{Cu} = [\text{Ar}] 4s^1 3d^{10}$

(C) As the shell no. inc. size of subshell increases e.g., size of  $(2s > 1s)$ ;  $(3p > 2p)$ ;  $(4d > 3d)$

(D) The region is an orbital for no probability of finding the  $e^-$  is known as **Nodal plane** (or Node).

The no. of [Spherical nodes or radial nodes] =  $n - l - 1$ .

(E) $\psi$ (psi)	$\psi^2$ (psi square)
A wave function for locating an $e^-$	The square of wave function where the probability of finding the $e^-$ is maximum. [Each value of $\psi^2$ is a region and <b>defines</b> one orbital]
(F) Orbit	Orbital
(1) A definite distance from the nucleus for finding the $e^-$ [ $e^-$ as a particle]	(1) A probability region for locating the $e^-$ around the nucleus. It is a wave function [ $e^-$ as a wave]
(2) It has definite size and $e^-$ in this orbit has definite energy.	(2) It does not define definite size. But only a boundary region diagram of a wave for locating the $e^-$ .

## STRUCTURE OF ATOM

### A. Fundamental particles of an atom, Electromagnetic Spectrum & Bohr's Theory

#### 1-Mark Questions

- Q. 1.** Give the difference in the origin of cathode and anode rays.
- Q. 2.** Which of the following will not show deflection from the path on passing through an electric field ? Proton, electron and neutron.
- Q. 3.** Neutron or proton of same wavelength, which will have more velocity ?
- Q. 4.** Chlorophyll absorbs radiations of wavelength 700 nm. What is the frequency of this radiation ? [Ans.  $4.2 \times 10^8$  Hz]
- Q. 5.** If the length of the crest of a wave is 4 pm. Write the wavelength of this wave. [Ans. 8 pm]
- Q. 6.** Arrange the following in the increasing order of their  $e/m$  values.  
electron, proton and neutron.
- Q. 7.** A radiation emitted from a hot iron is photon or quantum ?
- Q. 8.** Give two examples from everyday life where cathode rays tube used.
- Q. 9.** The line spectrum of an element is known as fingerprints of its atom. Comment
- Q. 10.** An element has atomic number 30 and mass number 66, what will be the number of protons and neutrons in this atom?  
[Ans. protons = 30 & neutrons = 36]
- Q. 11.** Calculate the mass and charge of one mole of electrons. [Ans.  $5.48 \times 10^{-4}$  g]
- Q. 12.** Calculate the charge of one mole of electrons. [Ans.  $9.6488 \times 10^4$  C]
- Q. 13.** What is the value of the Bohr's radius for the first orbit of hydrogen atom ?
- Q. 14.** Distinguish between a photon and a quantum,
- Q. 15.** What type of metals are used in photoelectric cell ? Give one example.  
[Ans. With large size less work function.]
- Q. 16.** Write the number of electrons present in 1 mol. of  $N^{3-}$  ion.
- Q. 17.** Write the difference between atomic mass and atomic mass number ?

- Q. 18.** Which series of lines of the hydrogen spectrum lie in the visible region ?
- Q. 19.** Cs show maximum photoelectric effect, why ?
- Q. 20.**  $\text{Cl}$  and  $\text{Cl}^-$  have same atomic number or mass number or both. Explain.

### 2-Mark questions

- Q. 1.** How could Rutherford in his experiment explain Nucleus ?
- Q. 2.** Define black body and black body radiations.
- Q. 3.** Give the essential postulates of Bohr's model of an atom. How did it explain ?
- (i) the stability of the atom ?
- (ii) origin of the spectral lines in H-atom ?
- Q. 3.** What is quantisation ? How quantisation of energy was introduced in Bohr's model ?
- Q. 4.** An element with mass number 81 contain 31.7% more neutron as compared to protons.

Assign the atomic symbol.

[Ans.  $^{81}\text{Br}$ ]

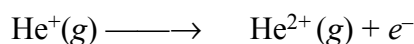
- Q. 5.** What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition  $n = 4$  to  $n = 2$  of  $\text{He}^+$  spectrum?

[Ans.  $n_1 = 1$  and  $n_2 = 2$ ]

- Q. 6.** What transition of  $\text{Li}^{2+}$  spectrum will have the same wavelength as that of the second line of Balmer series in  $\text{He}^+$  spectrum ?

[Ans.  $n_2 = 6$  to  $n_1 = 3$ ]

- Q. 7.** Calculate the energy required for the process



The ionization energy for the H atom in the ground state is  $2.18 \times 10^{-18} \text{ J atom}^{-1}$

[Ans.  $8.72 \times 10^{-18} \text{ J}$ ]

- Q. 8.** Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen. [Ans.  $1.523 \times 10^6 \text{ m}^{-1}$ ]

- Q. 9.** To which orbit the electron in H atom will jump on absorbing 12.1 eV energy ? [Ans. 3rd orbit]

- Q. 10.** Calculate the energy associated with the first orbit of  $\text{He}^+$ . What is the radius of this orbit. [Ans.  $-54.38 \text{ eV}$ ,  $0.2645 \text{ \AA}$ ]

- Q. 11.** What is the distance of separation between 3rd and 4th orbit of H-atom? [Ans. 3.703 Å]
- Q. 12.** The energy of electron in the first Bohr's orbit is  $-13.6$  eV. Calculate the energy of electron in the first excited state. [Ans.  $-3.4$  eV]
- Q. 13.** Calculate the number of photons emitted in 10 hours by a 60 W sodium lamp emitting radiations of wavelength 6000 Å.
- Q. 14.** Which one has a higher energy, a photon of violet light with wavelength 4000 Å or a photon of red light with wavelength 7000 Å ?  
[Given.  $h = 6.62 \times 10^{-34}$  J sec.]
- Q. 15.** A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.  
[Ans.  $2.012 \times 10^{20} \text{ s}^{-1}$ ]
- Q. 16.** What is the maximum number of emission lines when the excited electron of a H atom in  $n = 4$  drops to the ground state ? [Ans. 6]
- Q. 17.** Which has more energy, light radiation of wavelength 400 nm or light radiation of frequency  $10^{15}$  Hz ?
- Q. 18.** Find the energy of electron in 4th shell of  $\text{Li}^{2+}$  ion.
- Q. 19.** What is the wave number of an electron with shortest wavelength radiation in Lyman spectrum of  $\text{He}^+$  ion.
- Q. 20.** Write short note on :
- Continuous and discontinuous spectrum.
  - Absorption and emission spectrum.

### 3-Mark Questions

- Q. 1.** How were cathode rays discovered ? With the help of suitable experiments show that :
- Cathode rays travel in straight lines,
  - Cathode rays consist of material particles,
  - Cathode rays consist of negatively charged particles.
- Q. 2.** Write the complete symbol for :
- the nucleus with atomic number 56 and mass number 138.
  - the nucleus with atomic number 26 and mass number 55.

(iii) the nucleus with atomic number 4 and mass number 9.

**Q. 3.**(i) The energy associated with the first orbit in the hydrogen atom is  $-2.18 \times 10^{-18} \text{ J atom}^{-1}$ . What is the energy associated with the fourth orbit ?

(ii) Calculate the radius of Bohr's third orbit for hydrogen atom.

[Ans.  $-1.36 \times 10^{-19} \text{ J atom}^{-1}$ , 4.761 nm]

**Q. 4.** A bulb emits light of wave length 4500 Å. The bulb is rated as 150 watt and 8% of the energy is emitted as light. How many photons are emitted by the bulb per second ? [H.O.T.S1 [Ans.  $n = 27.2 \times 10^{18}$ ]

**Q. 5.** When light with a wavelength of 400 nm falls on the surface of sodium, electrons with a kinetic energy of  $1.05 \times 10^5 \text{ J mol}^{-1}$  are emitted.

(a) What is the minimum energy needed to remove an electron from sodium ?

(b) What is the maximum wavelength of light that will cause a photoelectron to be emitted ?

[Ans.  $a = 3.2255 \times 10^{19} \text{ J}$ ,  $b = 616 \text{ nm}$ ]

**Q. 6.** Compare the frequency of light radiations emitted when electron falls from 5th shell to the 2nd shell in  $\text{Li}^{2+}$  ion and electron falls from 4th shell to the 1st shell in  $\text{He}^+$  ion.

### 5- Mark Questions

**Q. 1.**(a) Define Photoelectric effect ? Mention its one practical application in daily life,

(b) Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 Å. Calculate threshold frequency ( $\nu_0$ ) and work function ( $W_0$ ) of the metal.

[Ans.  $\nu_0 = 4.41 \times 10^{14} \text{ s}^{-1}$ ,  $W_0 = 2.92 \times 10^{-19} \text{ J}$ ]

**Q. 2.**(a) The electronic energy in Bohr's orbit is negative. How will you account for it

(b) The ionisation energy of hydrogen atom is 13.6 eV. What will be the energy of the first orbit of  $\text{He}^+$  and  $\text{Li}^{2+}$  ions ?

[Ans.  $E_1$  of  $\text{He}^+ = -54.4 \text{ eV}$ ,  $E_1$  of  $\text{Li}^{2+} = -122.4 \text{ eV}$ ]

**Q. 3.**(a) Define the following terms :

(i) Threshold frequency

(ii) Work function.

- (b) The work function for Cs atom is 1.9 eV. Find threshold wavelength ( $\lambda_0$ ) and threshold frequency ( $\nu_0$ ) of this light radiation. If Cs metal is irradiated with a radiation of wavelength 500 nm find kinetic energy and velocity of emitted electron.

### Dual Nature of Matter, de Broglie Equation,

### Heisenberg's uncertainty principle and wave mechanical model

#### 1 - Mark Questions

- Q. 1.** Mention the physical significance of  $\psi$  and  $\psi^2$ .
- Q. 2.** Why did Heisenberg replace the concept of definite orbit by the concept of probability? ( $e^-$  with very less mass more as wave)
- Q. 3.** What is uncertain in uncertainty principle?
- Q. 4.** Can a moving cricket ball have a wave character? Justify your answer.
- Q. 5.** Heisenberg uncertainty principle has no significance in our everyday life. Explain.
- Q. 6.** Out of the  $d$  orbitals which does not have four lobes?
- Q. 7.** Write the Schrodinger wave equation.
- Q. 8.** Why uncertainty in position is more when uncertainty in velocity is less for an electron?

#### 2-Mark Questions

- Q. 1.** Calculate the mass of the photon with wavelength of 3.6 Å.  
[Ans.  $6.135 \times 10^{-29}$  kg]
- Q. 2.** Calculate the mass of the photon with wavelength of 5 Pm.
- Q. 3.** On the basis of uncertainty principle show that an electron cannot exist within atomic nucleus. (Given : Nuclear radius =  $10^{-15}$  m)  
[Hint : Taking  $10^{-15}$  m as  $\Delta x$ , the  $\Delta v$  comes much higher than the velocity of light and hence is not possible]
- Q. 4.** Explain why the uncertainty principle is significant only from the motion of subatomic particles and is negligible for macroscopic particles.
- Q. 5.** List two differences between orbit and orbital.
- Q. 6.** Show that the circumference of the Bohr orbit for the hydrogen atom



is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit

- Q.7. Comment on “Bohr’s model is against the Heisenberg uncertainty principle”.
- Q. 8. What are the similarities and difference in  $2s$  and  $2p_x$  orbitals and  $1s$  and  $2s$  orbitals ?
- Q. 9. Give three differences between  $3p$  and  $4p$  subshell.

### 3-Mark Questions

- Q. 1. Calculate the number of waves made by Bohr electron in one complete revolution in its third orbit. [H.O.T.S] [Ans. 3]
- Q. 2. What should be the ratio of velocities of  $\text{CH}_4$  and  $\text{O}_2$  molecules so that they are associated with de Broglie waves of equal wavelength ? [Ans. 2]
- Q. 3. Calculate the wavelength of an electron that has been accelerated in a particle accelerator through a potential difference of 1 kv.  
[Given  $1\text{eV} = 1.6 \times 10^{-19}\text{J}$ ] [H.O.T.S] [Ans.  $3.87 \times 10^{-7}\text{m}$ ]
- Q. 4.(i) Discuss the similarities and differences between a  $1s$  and  $2s$  orbital.  
(ii) Draw the shape of  $d_{z^2}$ .
- Q. 5. Calculate the wavelength of a tennis ball of mass 60 gm moving with a velocity of 10 m per second. [Ans.  $10^{-3}\text{metre}$ ]
- Q. 6. Calculate the wavelength of 1000 kg rocket moving with a velocity of 3000 Kilometre per hour. [Ans.  $7.9512 \times 10^{-40}\text{m}$ ]
- Q. 7. Calculate the uncertainty in the velocity of a cricket ball of mass 150 g, if uncertainty in its position is of the order of  $1\text{\AA}$ .  
[Ans.  $3.5 \times 10^{-24}\text{m s}^{-1}$ ]
- Q. 8.(a) What is de-Broglie wavelength for an electron with light velocity?  
(b) What is the angular momentum of electron in 5th shell.
- Q. 9. Two particles A and B have wavelength  $\lambda_A = 5 \times 10^{-10}\text{m}$  and  $\lambda_B = 10 \times 10^{-10}\text{m}$ . Find their frequency, wave number and energies. Which has more penetrating power and why ?
- Q. 10.(a) Which has max. uncertainty regarding position and why ?  
Electron, proton and neutron.

- (b) Find the number of waves associated with a light radiation of time period 5 ns.

**Q. 11.** If an electron in  $\text{He}^+$  has angular momentum of  $5h/2\pi$ . Find its energy and wavelength associated with it. Find the kinetic energy of this electron.

### 5-Mark Questions

**Q. 1.**(a) State de Broglie equation. Write its significance.

- (b) A beam of helium atoms moves with a velocity of  $2.0 \times 10^3 \text{ m s}^{-1}$ . Find the wavelength of the particle constituting the beam

$$(h = 6.626 \times 10^{-34} \text{ J s}) \text{ [Ans. } 49.9 \text{ pm ]}$$

**Q. 2.**(a) State Heisenberg's uncertainty principle. Give its mathematical expression. Also give its significance.

- (b) Calculate the uncertainty in the position of a dust particle with mass equal to 1 mg if the uncertainty in its velocity is  $5.5 \times 10^{-20} \text{ ms}^{-1}$ .

$$\text{[Ans. } 9.55 \times 10^{10} \text{ m]}$$

**Q. 3.**(a) Cricket ball, a tennis ball and a proton which has more uncertainty in velocity and which follows Heisenberg uncertainty principle maximum.

- (b) What is same in de-Broglie and Heisenberg principle, which is different from Bohr theory for structure of atom.

(c) Why Schrodinger wave only defines electron path uncertain but definite energy.

## Quantum Numbers, Electronic Configuration and Nodes

### 1 Mark Questions

**Q. 1.** What is the lowest value of  $n$  that allows  $g$  orbitals to exist ?

**Q. 2.** What are the four quantum numbers of 19<sup>th</sup> electron of copper ?

(Ans. Atomic number = 29)

**Q. 3.** Which quantum number is not obtained from solution of Schrodinger wave equation ?

**Q. 4.** How many electrons will be present in the sub-shells having  $m_s$  value of  $-1/2$  for  $n = 4$  ?

- Q. 5.** Write the electronic configuration of  $\text{Ni}^{2+}$ . (At. no. of Ni = 28)
- Q. 6.** How many radial and angular nodes are present in  $2p$  orbital.  
[Ans. Radial nodes = 0, Angular nodes = 1]
- Q. 7.** Which of the following orbitals are possible ?  
 $1p$ ,  $2s$ ,  $2p$  and  $3f$ .
- Q. 8.** Write the name of non-directional orbital.
- Q. 9.** Write the name of quantum numbers which determines the orientation of orbital and shape of orbitals.
- Q. 10.** Write the name of quantum number which determines the shape of orbital.
- Q. 11.** Using  $s, p, d$  notations, describe the orbital with the following quantum numbers :  
(a)  $n = 4, l = 2$  (b)  $n = 1, l = 0$ . [Ans. (a)  $4d$  (b)  $1s$ ]
- Q. 12.** How many orbitals are presenting subshell ?
- Q. 13.** How many total electrons can be filled in all orbitals with  $(n + l) = 5$  ? [Ans. 18 electrons ( $4p^6 3d^{10} 5s^2$ )]
- Q. 14.** Name the dipositive ion represented by the electronic configuration :  $1s^2 2s^2 2p^6 3s^2 3p^6$ . [ $\text{Ca}^{2+}$ ]
- Q. 15.** Is it correct to say that every atom with even atomic number has all electron paired ?
- Q. 16.** Cr in ground state has how many unpaired electrons.  
(Cr, Atomic number = 24).
- Q. 17.** Which has more energy of electron  $4p$  or  $5s$  ?
- Q. 18.** Nitrogen has correct configuration of  $1s^2, 2s^2, 2p_x^1, 2p_y^1, 2p_z^1$  is described by which principle ?
- Q. 19.** What are degenerate orbitals ?

### 2-Mark Questions

- Q. 1.** On the basis of Pauli's exclusion principle show that the maximum number of electrons in the M -shell ( $n = 3$ ) of any individual atom is 18.
- Q. 2.** Designate each subshell with  $n = 4$ .

- Q. 3.** List the possible values for all the quantum numbers for the following subshell.
- (a)  $2p$  (b)  $4f$
- Q. 4.** Write down the electronic configuration of  $\text{Fe}^{3+}$  and  $\text{Ni}^{2+}$ . How many unpaired electrons are present ? (Atomic number, Fe = 26, Ni = 28)
- Q. 5.** Out of principal, angular, magnetic and spin quantum number, which quantum number determines the ?
- (a) Shape of the orbital  
(b) Number of orbitals in an orbit  
(c) Size of the orbital  
(d) Spin orientation of the electron.
- Q. 6.** What is the Hund 's rule of maximum multiplicity ? Explain with suitable example.
- Q. 7.** Explain why :
- (a) The three electron present in  $2p$  subshell of nitrogen remain unpaired  
(b) Cr has configuration  $3d^5 4s^1$  and not  $3d^4 4s^2$ .
- Q. 8.** (a)  $4p^1$  and  $4p^2$  electron will have same (i)  $n, l, m$  or (ii)  $n, l, s$  or (iii)  $n, l, m, s$ , or  $n, l$   
(b) Nitrogen has 7 proton, 7 electron and 7 neutrons. Calculate the number of electron protons and neutrons in  $\text{N}^{3-}$  ion.
- Q. 9.** Which has more energy of electron :
- (a) Last electron of  $\text{Cl}^-$  or last electron of  $\text{O}^{2-}$ .  
(b)  $n = 4, l = 3$  or  $n = 5, l = 2$ .

### 3-Mark Questions

- Q. 1.**(i) An atomic orbital has  $n = 2$ . What are the possible values of  $l$  and  $m_l$  ?  
(ii) List the quantum numbers ( $m_l$  and  $l$ ) of electrons for  $3d$  orbital.  
(iii) Which of the following orbitals are possible ?  
 $2d, 1s, 2p$  and  $3f$ .
- Q. 2.**(a) Write the maximum number of electron in a subshell with  $l = 3$  and  $n = 4$ .

- (b) Write the maximum number of electron that can be associated with the following set of quantum numbers ?

$$[n = 3, l = 1 \text{ and } m_l = -1]$$

- (c) Write the maximum number of electron that can be accommodated in an atom in which the highest principal quantum number value is 4.

**Q. 3.**(i) Write the electronic configurations of the following ions :

(a)  $\text{H}^-$  (b)  $\text{Na}^+$  (c)  $\text{O}^{2-}$  (d)  $\text{F}^-$

- (ii) What are the atomic numbers of elements whose outermost electrons are represented by (a)  $3s^1$  (b)  $2p^3$  and (c)  $3p^5$  ?

- (iii) Which atoms are indicated by the following configurations ?

(a)  $[\text{He}] 2s^1$  (b)  $[\text{Ne}] 3s^2 3p^3$  (c)  $[\text{Ar}] 4s^2 3d^1$ .

**Q. 4.** Calculate:

- (a) Total number of spherical nodes in  $3p$  orbital.  
(b) Total number of Nodal planes in  $3p$  orbital.  
(c) What is nodal planes in  $3d$  orbital,

### 5-Mark Questions

**Q. 1.**(a) Write short notes on:

(i) Aufbau principle (ii) Pauli's principle (iii) Hund's rule.

- (b) Write the electronic configuration of the following ions :

(i)  $\text{Fe}^{3+}$  (ii)  $\text{Cu}^+$  [Atomic number of Fe and Cu are 26 & 29]

**Q. 2.**(a) Draw the shapes of the following orbitals.

(i)  $3d_{xy}$  (ii)  $dz^2$

- (b) What is the total number of orbitals associated with the principal quantum number  $n = 3$  ?

(c) Using  $s$ ,  $p$ ,  $d$ ,  $f$  notations, describe the orbital with the following quantum numbers.

(a)  $n = 3, l = 0$ , (b)  $n = 4, l = 2$ , (c)  $n = 5, l = 3$ , (d)  $n = 1, l = 0$

**Q. 3.** Explain the following :

- (i) Energy of electron is not decided by :  $n, l, m$  and  $s$ .

- (ii) Maximum number of electron with  $-1/2$  spin for  $n = 3$  is 6,9,12 or none.
- (iii) Maximum number of electron can be present for  $n + l = 4$ .
- (iv)  $3f$  subshell is not possible.
- (v) Maximum number of electrons in a subshell is :  
 $(2l + 1)$  or  $(4l + 1)$  or  $n^2$

**Q. 4.**(a) A neutral atom has 2K, 8L & 15 M electrons. Find the total numbers of electrons in  $s$ ,  $p$ ,  $d$  and  $f$  subshell.

- (b) How many unpaired electrons are present in the following ions :

$\text{Al}^+$ ,  $\text{Cr}^{2+}$ ,  $\text{Co}^{3+}$  and  $\text{Mn}^{2+}$

(Atomic number :  $\text{Al}=13$ ,  $\text{Cr} = 24$ ,  $\text{Co} = 27$  &  $\text{Mn} = 25$ )

- (c) One electron is present in  $4f$  subshell. What is the sum of  $n + l + m_1 + m_s$  values assuming ' $f$ ' subshell follows  $-3$  to  $+3$  order of filling electron.

**Q. 5.** Write the :

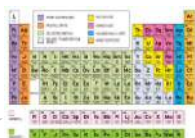
- (a)  $n + l$  value for  $14^{\text{th}}$  electron in an atom.
- (b) Increasing order of filling electron in  $4f$ ,  $5p$  and  $6d$  subshells.
- (c) ' $m$ ' and ' $l$ ' value for last electron of Mg atom.

(Mg atomic number is 12)

- (d) Subshell in which last electron is present in Ga.

(Atomic number is 31)

- (e) Sum of spin of all the electron in atomic number 14.



## Chapter - 3

# Classification Of Elements And Periodicity In Properties

The first classification of elements was provided by Russian chemist D.I. Mendeleev.

1. “The physical and chemical properties of elements are periodic functions of their atomic weight.”
2. It was modified to **Modern Periodic law** : “The physical and chemical properties of elements are periodic functions of their atomic numbers.”

It is the long form of periodic table :

Horizontal rows → Periods

Vertical columns → Group

3. 1<sup>st</sup> period—2 elements  
2<sup>nd</sup> and 3<sup>rd</sup> period—8 elements  
4<sup>th</sup> and 5<sup>th</sup> period—18 elements  
6<sup>th</sup> period—32 elements  
7<sup>th</sup> period—Incomplete
4. Groups  
1 and 2 — ‘s’ block elements last electron entered in ‘s’ subshell [ $s^1, s^2$ ]  
3 to 12 — ‘d’ block elements last electrons entered in ‘d’ subshell [ $d^1$  to  $d^{10}$ ].  
13 to 18 — ‘p’ block elements last electrons enter in ‘p’ subshell [ $p^1$  to  $p^6$ ].  
Group 18 — Noble gases.
5. (A) In ‘s’ and ‘p’ block elements the electrons enters outer most shell.  
In ‘d’ block elements the electron enter the penultimate shell ( $n - 1$ ).  
‘f’ block elements last electron enter the subpenultimate shell ( $n - 2$ ).  
(B) ‘f’ block elements are placed in between ‘d’ block elements.  
‘f’ block elements in 2 rows [4f lanthanoids 5f actinoids]  
(C) Helium is placed  $ns^2$ . But it has Noble gas configuration.  
(6) General electronic configuration ‘s’ block  $ns^1, ns^2$  [Group I]

'p' block  $ns^1 np^1$  to  $ns^1 np^6$  Group 13 to 18

'd' block  $ns^{1,2} (n-1) d^1$  to  $10$  Group 3 to 12

'f' block  $ns^2 (n-1) d^{0,1} (n-2) f^1$  to  $14$

## 7. General Trends :

### Atomic Radius

- (A) Left to right decreases due to effect of successive increasing nuclear charge without addition of a new shell.
- (B) From top to bottom atomic radius increases due to successive addition of shell.
- (C) Noble gases large radius than **group 17** due to complete filling of electron in outer shell electron-electron repulsion mildly increases.

**8. (A) Ionization energy :** The energy required to remove electron from outer most shell of an atom in its isolated gaseous state.  $M_{(g)} \longrightarrow M_{(g)}^+ + e^-$

**(B) From Left to right** increases due to decrease in atomic radius therefore less force of attraction by protons. For 2<sup>nd</sup> group elements.

1<sup>st</sup> ionization enthalpy

$Li < B < Be < C < O < N < F < Ar$

II<sup>nd</sup> ionization enthalpy

$Be < C < B < N < F < O < Ne < Ni$

**(C) From top to bottom :** Due to increase in size, (due to addition of shells) the ionization enthalpy decreases from top to bottom.

**(D) Metallic behaviour :** Decrease from left to right due to increase in ionization enthalpy.

**Non metallic behaviour :** Increase from left to right due to more number of electron in outershell and added electron goes towards nucleus.

**(A) Ionic radius :** (A) Cation radius < Atomic radius—due to more no. of protons than number of electron coulombic force increases size decreases.

$[Mg^{2+} < Mg^{+1} < Mg]$

**(B) Anion radius > Atomic radius—**Due to more number of electron than number of protons.  $[N^{3-} > O^{2-} > F^-]$

Electron-Electron repulsion increases, coulombic force of attraction decreases.

**(C) For Isoelectronic species—**More is the charge of cation lesser in the size.



More is the change of anion more is the size.

(D) Size— $O^{2-} > F^- > Na > Na^+ > Mg^{2+}$

(10) **Electro negativity** : (A) The tendency of an atom to attract the shared pair of electron towards itself in a bonded state.

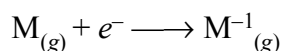
(B) Left to right size decreases shared electron nearer to nucleus—electron negativity increases.

(C) Top to bottom size increases, shared electron away from nucleus—electron negativity decreases.

(D) Four highest electronegative atoms  $F > O > Cl > N$

**maximum electronegative Assign to F.**

(11) **Electron gain enthalpy** : The energy rescaled when an extra electron is added to neutral gaseous state.



**Negative electron gain enthalpy**

(A) **Trends—From left to right**—Increase due to decrease in size more attraction of added electron by nucleus.

(B) **From top to bottom**—Decreases as the added to electron is away from nucleus due to increase in size.

(C) **Cl has more electron gain enthalpy than fluorine**—Due to small size of fluorine extra added electron has more inter electronic repulsion than chlorine which has large size. Similarly **Phosphorus and Sulphur** have negative electron gain enthalpy than **nitrogen and oxygen** respectively.

(D) **Maximum electron gain enthalpy**—Chlorine (in periodic table)

(E) Electron gain enthalpy—

Halogen > Oxygen > Nitrogen > Metal of group 1 and 13 and non metal of group 14 > metal of group 2.

(F) II electron gain enthalpy in always five.

**Electron affinity**—The negative of the enthalpy change accompanying the addition of an electron to an isolated gaseous atom is defined as **electron affinity** (Ae).

+ **ive electron affinity**—If energy released when an isolated gaseous atom accepts an electron.

– **ive electron affinity**—If energy is to be supplied to add an extra electron

to the isolated gaseous atom.

**(12) Screening effect—(A)** The decrease in force of attraction exerted by nucleus on valency electrons due to presence of electrons in inner shells is known as screening or shielding effect.

**(B)** Effective nuclear charge ( $Z^*$ ) = Nuclear charge—Shielding effect

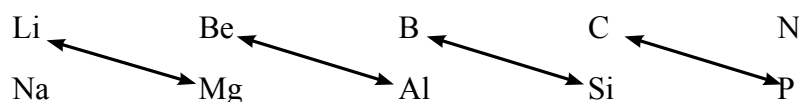
**(C)** Trend—Left to Right— $Z^*$  Increases.

Top to Bottom— $Z^*$  Decreases.

**(13) Second period element—**Show different behaviour that I group element—Due to (a) small size (b) High electron negativity (C) High polarising power (d) absence of 'd' orbital in I member.

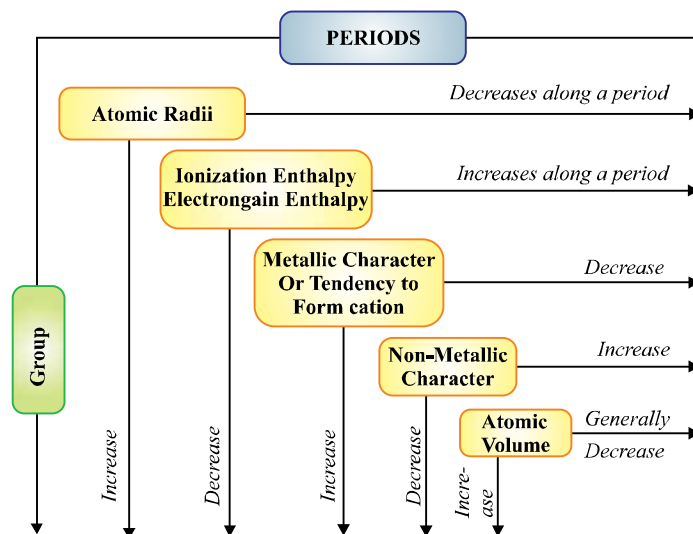
$\text{Na}_3[\text{Al}(\text{OH})_6]$  exists but  $\text{Na}[\text{B}(\text{OH})_4]$  not exists.

**(14)** The similarities in properties of I member of a group to II member of just next higher group due to comparable atomic radius, nearly same polarising power of ions is known as **Diagonal relationship**.



**Elements with outer shell**

- |                    |            |
|--------------------|------------|
| (a) 1, 2, 3, $e^-$ | metals     |
| (b) $4e^-$         | metalloids |
| (c) 5, 6, 7, $e^-$ | non-metals |
| (d) $8e^-$         | noble gas  |



## Arrangement of Elements

### 1-Mark Questions

1. Metals are placed on which side of modern periodic table ?
2. Which pair of elements has similar properties ?  $\begin{bmatrix} 13, & 31 \\ 11, & 21 \end{bmatrix}$  [Ans. 13, 31]
3. Which atomic number refers to a non-metal elements ? [Ans. 17, 35]
4. Eka Aluminium & Eka silicon are known as ..... and ..... .
5. (a) Be shows diagonal relationship with which element. (Gr 16) (Al)  
(b) Which group elements are known as chalcogens.
6. Which pair of elements are 'S' block elements ?  $\begin{bmatrix} 37, & 38 \\ 55, & 37 \end{bmatrix}$  [Ans. Both]
7. The element with  $ns^2, np^5$  configuration is non-metal or metal ?
8. Define Vanderwall radius. [The half of the distance b/w two nearly place atoms in gaseous state].
9. Write the outer shell configuration of atomic number 31. [Ans.  $4s^2 4p^1$ ]
10. Find the group no. and period no. atomic number 52. [Ans. 5<sup>th</sup> Row, 16 gr]
11. Arrange  $O^{2-}$   $O^{-1}$  O in decreasing radius (size). [Ans.  $O^{2-} > O^{-1} > O$ ]
12. Which element is isoelectronic with  $Na^+$  (Na atomic number 10)  
[Ans. Ne]
13. Why noble gas have more size than halogens (Gr 1 elements).
14. An element is placed in 5th row and 3 group has what atomic no. ?  
[Ans. 39]
15. How electron gain enthalpy and sulphur is more than  $O^-$  ?
16. If 'f' block elements 4f elements have what electronic configuration ?  
[Ans.  $6s^2, 5f^{0,1} 4f^{1 \text{ to } 14}$ ]
17. Write the IUPAC name of element with atomic number 115, 107.
18. Write two isoelectronic species with Br (35). [Ans.  $Kr^+$   $Se^{-1}$ ]
19. What is covalency of Al in  $[AlCl_4]^{1-}$  ? [Ans. 4]
20. Modern periodic table which groups have less no. of elements in modern period table. [Ans. 3 to 12]
21. Show that 4th period can have maximum 18 elements.

### Ionization energy electronegativity Electron Gain Enthalpy

22. 2<sup>nd</sup> IE is always more than first why ?
23. Electronegativity of  $F < Cl < Br < I$  why ?
24. Arrange F, Cl in terms of increasing chemical reactivity.
25. 2<sup>nd</sup> IE of Na is more than 2<sup>nd</sup> IE of Mg. Why ?
26.  $\Delta_{eg}H^\ominus_1$  is exothermic while  $\Delta_{eg}H^\ominus_2$  is endothermic justify.
27. IE for cation is more than for neutral atom. Why ?
28. Arrange in decreasing order of first electron gain enthalpy exothermic .  
[Ans. N, O]
29.  $O^{1-}$  O which has exothermic and which has endothermic electron gain enthalpy.

### Trends, IE, Electronegativity and Electron gain enthalpy Metallic behaviour valency

#### 2-Mark Questions

1. Cations are smaller than neutral atom why anions are larger in size than neutral atom ?
2. Ionization energy of nitrogen is more than 'O' and 'C' both why ?
3. First ionization energy of boron is less than Be but size of Be is less than Boron. Why ?
4. Electron gain enthalpy of Mg is positive why ?
5. The covalency of Si is 4 but of halogens is one why ?
6. The reactivity of halogens decrease down the group but of alkali metals (Gr 1) inc. down the group.
7. Name one halogen, one metal one gr 13 element which are liquid at 30°C.  
[Ans. Br, Hg, Ga]
8. The reducing power of elements increases down the group but reverse is true for oxidising power along a period. Why ?
9. What is formula of binary compound formed b/w combination of :  
(a) 1<sup>st</sup> element of 1<sup>st</sup> group and iodine.  
(b) 2<sup>nd</sup> element of 2<sup>nd</sup> group and 1<sup>st</sup> element of 17 groups.
10. Arrange in increasing order of :  
Size (a) I, F, Cl, Br

Oxidising power (b) I, F, Br, Cl

11. Oxygen is more metallic than nitrogen but less than fluorine why ?
12. LiCl, LiBr, LiI are covalent as well as ionic why ?
13.  $\text{PbCl}_2$  is more stable than  $\text{PbCl}_4$ . Why ? [Ans. Inert pair effect]
14. [Magnesium and nitrogen] and [Lithium and nitrogen] both form nitrides why ?
15. Which has least IE [ $3p^3$ ,  $3p^6$ ,  $2p^3$ ,  $2p^6$ ]
16. (a) IE of sulphur is lower than chlorine.  
(b) Arrange in decreasing electronegativity F, O, N, Cl, C, H.
17. Element 'A' in Gr 17 2<sup>nd</sup> row  
'B' in Gr 16 2<sup>nd</sup> row  
'C' in Gr 15 2<sup>nd</sup> row  
Arrange 'A', 'B' and 'C' in decreasing EN, IE and 1<sup>st</sup> EGE.
18. Element 'A' 13 group forms ionic compounds. Write the :  
(a) Formula of its oxide.  
(b) Arrange of decreasing electro + ve character Ga, Li, Pb, Cs.
19. Write the atomic no. of element placed diagonally to element of :  
(a) Group 14 row 4 (b) Group 2 Row 5  
(c) Group 17 row 4
20. An element has outer shell configuration  $4s^2 4p^3$ .  
(a) Find the atomic number of element placed below it.  
(b) Next noble gas no.

### 3-Mark Questions

1. What is metallic radius, Covalent radius, Vanderwall radius. Give one example for each.
2. Oxygen has first electron gain enthalpy exothermic second endothermic still a large no. of ionic oxides are formed. Why ?
3. Boron and Al give two similarities and two behaviour shown differently by each.
4. Element Gr 17; Gr 18; Gr 1 in the sequence placed in modern periodic table.  
(a) Which has – ve electron gain enthalpy ?

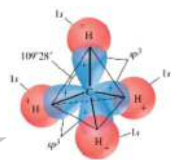
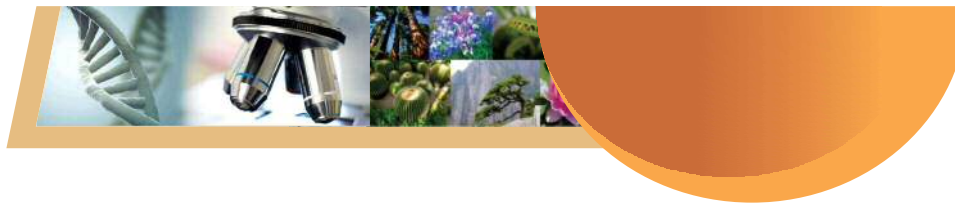
- (b) Which has more metallic behaviour ?  
 (c) Which has zero electronic behaviour ?
5. What are (a) representative elements, (b) Transition elements, (c) Lanthanoid and actinoids. Give their positions in modern periodic table.
6. Why LiF NaF KF RbF CsF are ionic ? But LiF is less ionic than CsF.
7. (a) Why Ca has atomic radius than Al ?  
 (b) Why  $2s^2 e^-$  is difficult to remove than  $2p$  electron ?
8. (a) The non-metallic behaviour is shown by halogen (Gr 17) but their compounds with (Gr 13) elements are more ionic and stable than with (group 1) elements. (b)  $Al_2O_3$  is more ionic than MgO than  $Na_2O$ .
9. Explain the following data :  
 Ionization energy  $Cl < H < O < N < F$ .
10.  $IE_2$  of 3<sup>rd</sup> period elements is as follows. Why ?  
 $Mg < Si < Al < P < S < Cl < Ar < Na$ .

### 5-Mark Questions

1. (A) Which of the following have same chemical properties :  
 (a) Atomic no. 17, 53 (b) Atomic no. 8, 52  
 (c) Both (d) None  
 (B) Which of the following more (IE)  
 (a) B, Al, Ga arrange in decreasing order.  
 (C) C, S, N first electron gain enthalpy (decreasing order)  
 (D) Al forms amphoteric oxide. Why ?  
 (E) Si is a semiconductor 'C' is a non-metal why ?

2. Element	$\Delta_f H^\ominus_1$	$\Delta_f H^\ominus_2$	$\Delta_{eg} H^\ominus_1$
I	1681	3374	- 328
II	1008	1846	- 295
III	2372	5251	+ 48

- (a) The most reactive non-metal.  
 (b) The least reactive non-metal.  
 (c) The least reactive element give reason also. [Ans. (a) I (b) II (c) III]

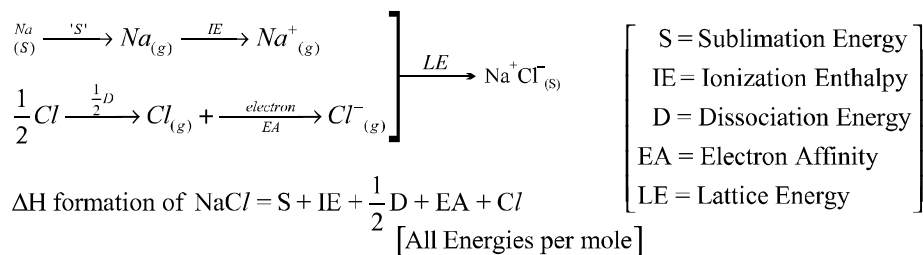


## Chapter - 4

# Chemical Bonding And Molecular Structure

1. The interaction between two atoms which holds them together within a molecule or ions is known as chemical bond.
2. The elements with one, two, three, four, five, six or seven electrons in outer shell, use these electrons to complete octet. The electrons which take part in two or more atoms to complete octet is known as electrovalency.
3. Lewis symbols or electron dot symbols involve the presentation of valence electrons (outer electrons) in an atom  $\dot{\text{Na}}$ ,  $\text{Mg}:$ ,  $\ddot{\text{Al}}$ ,  $\ddot{\text{Si}}$  etc.
4. Electrovalent bond or ionic Bond : The bond (chemical interaction) between two atoms formed by complete transference of electron from valence shell (outer shell) of an atom to another to complete octet (noble gas configuration)  $[2e^- \text{ in H, Li}]$  is known as ionic bond.
5. This ionic bond is favoured by low ionization enthalpy of metal high electron gain enthalpy of non-metal atom and in the resulting ionic compound more lattice energy.
6. Characteristics of ionic compound : They are solids, a definite arrangement/pattern of ion (to give crystalline solids), high MP and BP, conductors in fused state and in aqueous medium, soluble in  $\text{H}_2\text{O}$  [Hydration].
7. **Lattice enthalpy** : The energy released when one mole of ionic compound is formed from its ions in their gaseous state. Lattice energy is directly proportional to charge of ion and inversely proportional to size of ions *i.e.*, more is charge density more is lattice energy.
8.  $e.g., \text{Mg} \longrightarrow \text{Mg}^{2+} + 2e^- \quad \text{O} + 2e^- \longrightarrow \text{O}^{2-}$   
 $(2, 8, 2) \quad (2, 6)$   
 $\text{Mg}^{2+} + \text{O}^{2-} \longrightarrow \text{MgO}$  Ionic compound [A crystalline lattice].

9. (a) Born Haber Cycle : For formation of ionic compound e.g.,  $\text{Na}^+\text{Cl}^-$ .



(b) Ionic bonds are **Non directional** in nature.

10. **Fajan's Rule** : Polarizability and polarizing power. The power of an ion to distort the other ion is polarizing power and the tendency of an ion to get distorted is known as polarizability. Factors affecting polarizing power and polarizability.

(a) High charge and small size of  $\text{C}^+$ .

(b) High charge and large size of  $\text{A}^-$ .

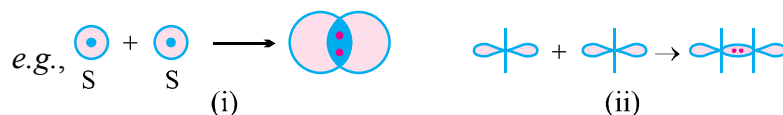
### 11. Covalent Bond : Lewis Langmuir Concept

The (chemical interaction) bond formed between two atoms say mutual sharing of electrons between them so as to complete their octets is known as covalent bond and no. of electrons involved is covalency.

12. **Formal charge** : Electron of an atom in a molecule/ion

$$\text{FC} = [\text{Total no. of valence in free atom}] - [\text{Total no. of non bonding electrons}] - \frac{1}{2} [\text{Total no. of shared electrons}]$$

13. **The valence bond approach** : (a) The two atomic orbitals with one electron each, overlap to give maximum electron density region common to both atoms is known as single covalent bond



(b) The strength of covalent bond is proportional to extent of overlapping between the atomic orbitals of valence shell.

14. Characteristics of covalent compounds—(a) They are in all three physical states solid, liquid, or gas depending upon factors like molecular mass, Vander-wall force, covalency, ( $S_8$ ), Hydrogen bonding, polarity etc.
- (b) These are directional soluble in less polar or non-polar solvents, less



MP and BP [than ionic compounds] Due to weak Vander-Waal forces, bad conductor of electricity (no. free electrons)

**15. Hybridization :** (a) Phenomenon of intermixing of atomic orbitals with slightly different energies to form new orbitals of equal energy and identical shape. The new orbitals are known as **hybrid orbitals**.

(b) The number of hybrid orbitals is equal to no. of orbitals mixing.

(c) As the 's' character in hybrid orbital inc. electronegativity and size of hybrid orbital inc.

(d) The hybrid orbital can have paired  $e^-$  as well as unpaired  $e^-$ .

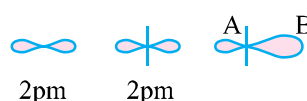
**16. VSEPR** (Valence shell electron pair repulsion theory)

(a) There are three types of repulsion in a covalent molecule  $lp - lp > lp - bp > bp - bp$ .

(b) These repulsions depend upon difference in electronegativity between A and B.

(c) The hybridization of central atom decides geometry and VSEPR theory the shape of molecule.

**17. Sigma and Pi  $\pi$  bond :** (a) The bond formed by overlap of two atomic orbitals along the internuclear axis of two atoms is Sigma bond.



Degree of overlap is **large therefore strong bond**.

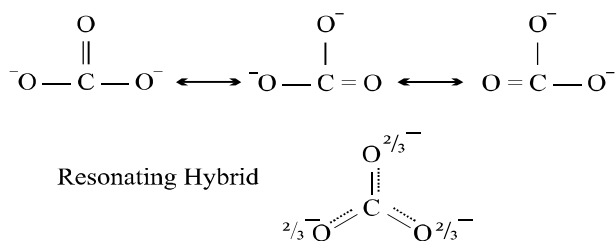
(b)  **$\pi$  bond :** The bond formed by sideways overlap of two atomic orbitals degree of overlap is not along the nuclear axis therefore a weak bond.

(c) A single bond is  $\sigma$  bond; A double bond is one  $\sigma$  one  $\pi$  bond a triple covalent bond is one  $\sigma$  and  $2\pi$  bonds.

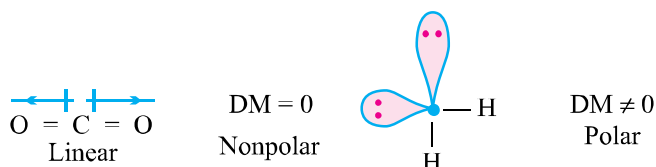
**18. Resonance :** (a) The delocalization of  $2e^-$  in a molecule/ion which results in observed bond length, bond order, bond energy different from normal covalent bond. Data is known as resonance. Various resonating structures have nearly same energy and interconvertible to each other.

(b) It gives stability to the molecule/ion. Atom do not shift their position in any of the resonating structure. The structure which is near to all resonating structure and nearly explain the property of that molecule/ion is known

as **resonating hybrid**.

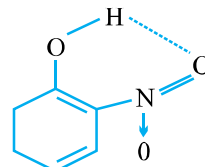


**19. Dipole moment :** (a) For polar covalent molecules (atoms with difference in electronegativity] the product of charge separation and distance b/w charges is known as dipole moment. (b) Being vector, if net resultant of all vector is zero the molecule has zero DM and known as non **polar** otherwise.

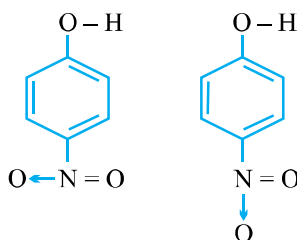


**20. Hydrogen bonding :** The dipole interaction b/w molecules when H is bounded with highly electronegative atoms (F, O, N only).

(a) **Intramolecular HB :** When hydrogen bonding is in between the same molecule. Hence molecules are independent and with less MP, BP. Due to within hydrogen bonding notable to make hydrogen bonding with  $\text{H}_2\text{O}$  therefore less soluble in water.



(b) **Intermolecular hydrogen bonding :** When hydrogen bonding is between the different molecules therefore close packing of molecules therefore more MP and BP and more soluble in water.



**21. Molecular orbital theory :** (a) The overlap of atomic orbitals of same symmetry of two homonuclear atoms to give addition or subtraction of wave functions and form bonding MO and antibonding MO respectively is known as MO theory.

(b) The  $e^-$  are filled in molecule increasing energy order of MO.

(c) **Bond order** : Bond order of molecule/ion

$$= \frac{\text{No. of electron in BMO} - \text{No. of electrons in ABMO}}{2}$$

(d) More is bond order more is bond energy lesser is the bond length.  
Bond order zero means no. possibility of that molecule.

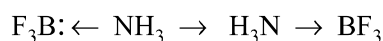
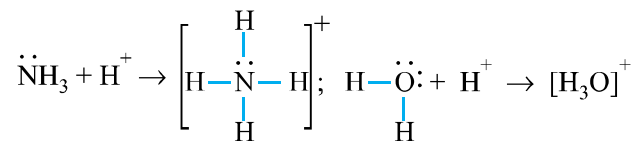
(e) Increasing order of energy of MO for upto  $14e^-$ .

$$\sigma_{1s}^*, \sigma_{1s}^*, \sigma_{2s}^*, \sigma_{2s}^*, \pi_{2px} = \pi_{2py}, \sigma_{2pz}$$

$$\text{for more than } 14e^- \rightarrow \sigma_{1s}^*, \sigma_{1s}^*, \sigma_{2s}^*, \sigma_{2s}^*, \sigma_{2pz}^*, \pi_{2px} = \pi_{2py}^*, \pi_{2px}^* = \pi_{2py}^*, \sigma_{2pz}^*$$

(d) This theory decides the magnetic behaviour also. [Equal energy orbitals]

**22. Coordinate covalent bond** : The sigma bond formed by donation of lp into vacant by drized orbital of other atom (acception atom) is known as coordinte covalent bond or **donor acceptor** or **daive bond**.



**23. Bond strength; Bond energy; Bond length :**

For covalent molecule BE  $\text{A} \equiv \text{A} > \text{B} = \text{B} > \text{C} - \text{C}$

BL  $\text{A} \equiv \text{A} < \text{B} = \text{B} < \text{C} - \text{C}$

**Bond angles** : (a) As the no. of  $lp$  increase;  $bp - lp$ ,  $lp - lp$  repulsion increase therefore bond angles decrease.

(b) As  $lp$  decreases bond angle increases.

**Important Points :**

**24.** (a) Max. density of  $\text{H}_2\text{O}$  at  $4^\circ\text{C}$  due to strongest.

Intermolecular hydrogen bonding.

(b)  $\text{H}_2\text{O}$  to  $\text{H}_2\text{S}$  bond angle decrease

(c)  $\text{NH}_3$  more DM than  $\text{NF}_3$ .

(d)  $\text{NCl}_3$  more DM than  $\text{NF}_3$ .

**25. Vander wall forces** weak intermolecular force of attraction.

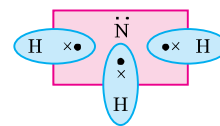
- (a) Dipole interactions  $\text{NH}_3$ ,  $\text{HCl}$ ,  $\text{HF}$ .
- (b) Dipole induced dipole interactions. Noble gas in  $\text{H}_2\text{O}$ ,  $\text{HCl}$ ;  $\text{F}_2$
- (c) Induced dipole induced dipole interaction (london force) Halogens noble gases.
- (d) Ion reduced dipole interaction  $[\text{NO}_3^- \text{Br}_2]$ .

## Chemical Bonding and Molecular Structure

### 1-Mark Questions

1. What is the covalence of Al in  $\text{AlCl}_3$  ? [Ans. 3]
2.  $\text{MgCl}_2$  is covalent or ionic compounds. [Ans. Both]
3.  $\text{CCl}_4$   $\text{CHCl}_3$   $\text{CH}_2\text{Cl}_2$  which is covalent. [Ans. All]
4.  $\text{Al}_2\text{O}_3$  has ionic or covalent bond. [Ans. Ionic]
5.  $\text{AlF}_3$  or  $\text{AlCl}_3$  which is covalent ? [Ans.  $\text{AlCl}_3$ ]
6. The  $[\text{BF}_4]^{1-}$  has what covalence ? [Ans. Four]
7. Ionic bonds are formed by metal with non-metals ? (True/False) [Ans. True]
8.  $\text{NH}_3$  and  $\text{NH}_4^+$  have what cordencies. [Ans. 3, 4]
9. Covalency can be mutual sharing or one sided donation of electrons (True/False) [Ans. True]
10.  $\text{N}_2$ ,  $\text{O}_2$ ,  $\text{H}_2$  have what covalencies. [Ans. 3, 2, 1]
11.  $\text{I}_2$ ,  $\text{Cl}_2$ ,  $\text{Br}_2$ ,  $\text{F}_2$  all have single covalent bond. (True/False) [Ans. True]
12.  $\text{AlCl}_3$  and  $\text{AlCl}_4^{-1}$   $\text{Al}_2\text{Cl}_6$  have what covalencies. [Ans. 3, 4, 4]
13.  $\text{C}_2\text{H}_6$  has 'C' with what covalencies between C and C. [Ans. 2]
14.  $\text{N}_2\text{H}_4$  (Hydrazine) has what covalency of N—N. [Ans. 1]
15.  $\text{HF}$   $\text{HCl}$   $\text{HBr}$  all polar covalent or pure covalent. [Ans. Polar]
16. Lewis dot structure of  $:\ddot{\text{N}} + 3e^- \longrightarrow \dots\dots$  [Ans.  $[\text{N}]^{3-}$ ]
17. Lewis dot structure of  $\text{AlN}$  is  $\longrightarrow \dots\dots$  [Ans.  $[\text{Al}]^3[\text{N}]^{3-}$ ]
18. Pure covalent diatomic molecule is  $\text{H}_2$ ,  $\text{C}_2$ ,  $\text{N}_2$ , all. [Ans. All]

19. Lewis dot structure of  $\text{NH}_3$  is  $\longrightarrow$  .....



20.  $\text{C}_6\text{H}_6$  has C with covalency of 4, 3, 2.

## Resonance; MO theory; Dipole moment

### 1-Mark Questions

- More is the no. of resonating structures more is the stability of that molecule/ion. (True/False) [Ans. True]
- Resonance involves shifting of  $\pi$  electrons with lp, negative charge, positive charge or all. [Ans. All]
- $\text{N}_2$  to  $\text{N}_2^+$  bond length increases. (True/False) [Ans. True]
- $\text{O}_2$ ,  $\text{O}_2^+$ ,  $\text{O}_2^-$  what is decreasing bond order. [Ans. ( $\text{O}_2^+ > \text{O}_2 > \text{O}_2^-$ )]
- $\text{H}_2^+$   $\text{H}_2$  which is more stable. [Ans.  $\text{H}_2$ ]
- More is the bond order more is the bond length between atoms. (True/False) [Ans. False]
- Dipole moment of  $\text{BeCl}_2 = \text{O}$ . (True/False) [Ans. True]
- All linear molecules have zero dipole moment. (True/False) [Ans. False]
- Dipole moment of a molecule/ion decides the polarity of that molecule/ion. (True/False) [Ans. True]
- $\text{CO}_2$ ,  $\text{BeCl}_2$ ,  $\text{ICl}_4^-$ ,  $\text{SF}_6$  which has zero DM ? [Ans. All]
- No. of antibonding electrons in  $\text{O}_2^+$  is 1, 2 or 3. [Ans. 1]
- $\text{N}_2$ ,  $\text{O}_2$ ,  $\text{F}_2$  arrange in decreasing order of stability. [Ans.  $\text{N}_2 > \text{O}_2 > \text{F}_2$ ]
- $\text{SO}_2$ ,  $\text{CO}_2$  which has zero dipole moment. [Ans.  $\text{CO}_2$ ]
- Dipole moment of hydrogen halides from HF to HI. Why ?  
[Ans. At radius increases EN Dec.]
- Which does not show resonance  $\text{CO}_3^{2-}$ ,  $\text{BO}_3^{3-}$ ,  $\text{SO}_4^{2-}$  ? [Ans.  $\text{BO}_3^{3-}$ ]

## VSEPR, Shape, Bond Angle, Bond Energy, Hybridization

### 2-Mark Questions

- Which are isostructural species [same shape]  $\text{NO}_2^+$ ,  $\text{CO}_2$ ,  $\text{BeCl}_2$ ,  $\text{BCl}_3$  ?  
[Ans.  $\text{NO}_2^+$ ;  $\text{CO}_2$ ,  $\text{BeCl}_2$  all linear]

2.  $\text{BCl}_3$ ,  $\text{H}_2\text{O}$ ,  $\text{NO}_3^-$  which have same type of hybridization. [Ans.  $\text{BCl}_3$   $\text{NO}_3^-$ ]
3. Which is angular or bent  $\text{XeF}_2$ ,  $\text{H}_2\text{O}$ ,  $\text{NO}_2$  ? [Ans.  $\text{XeF}_2$ ,  $\text{NO}_2$ ]
4. Which is not linear  $\text{XeF}_2$ ,  $\text{NO}_2^+$ ,  $\text{CO}_3^{2-}$ ,  $\text{ICl}_2^-$  ? [Ans.  $\text{ICl}_2^-$   $\text{XeF}_2$ ,  $\text{NO}_2^+$ ]
5. Which has more EN of carbon ?  
(a)  $\text{CH} \equiv \text{CH}$  (b)  $\text{CH}_2 = \text{CH}_2$  (c)  $\text{CH}_3 - \text{CH}_3$ .  
[Ans.  $a > b > c$  as 's' character increase EN increases]
6.  $\text{N}(\text{SiH}_3)_3$   $\text{N}(\text{CH}_3)_3$  are not isostructural why ?  
[Ans. Si has vacant 'd' orbital 'C' does not]
7. Which has maximum bond angle ?  
 $\text{CH}_4$ ,  $\text{BeCl}_2$ ,  $\text{NH}_3$  arrange in decreasing order. [Ans.  $\text{BeCl}_2 > \text{CH}_4 > \text{NH}_3$ ]
8. Hydrogen bond is shorter than  $\text{H}-\text{H}$  bond. (True/False) [Ans. True]
9. Which has nearly same bond angle ?  
 $\text{NH}_4^+$ ,  $\text{CCl}_4$ ,  $\text{CH}_3^+$  [Ans.  $\text{NH}_4^+$ ,  $\text{CCl}_4$ ]
10. Which has different hybridization ?  
(a)  $\text{AlCl}_3$ ;  $[\text{AlCl}_4]^-$  (b)  $\text{BF}_3$   $[\text{BF}_4]^-$   
(c)  $[\text{NH}_3 \text{NH}_4]^+$  (d)  $\text{NO}_2^+$  ( $\text{NO}_2^-$ ) [Ans. (c)]
11.  $\text{KHF}_2$  exist but  $\text{KCl}_2$   $\text{KBr}_2$  does not why ?  
[Ans. ( $\text{HF} \cdots \text{HF}$ ) Hydrogen bonding]
12. As EN increases polarity increases but still chlorine does not show hydrogen bonding but nitrogen shows (EN of Cl 3.2; N 3.0).  
[Ans.  $\because$  Cl has large size  $\text{H}-\text{Cl}$  is less polar]
13. Which is  $sp^3$  hybridized Be atom ?  
(a)  $\text{BeCl}_2(\text{g})$  (b)  $\text{BeCl}_2$  (solid) (c)  $\text{BeCl}_4^{2-}$ . [Ans.  $\text{BeCl}_2$  solid  $[\text{BeCl}_4]^{2-}$ ]
14.  $[\text{AlF}_6]^{3-}$   $[\text{Al}_2\text{O}_3]$  which have both covalent and ionic bond ? [Ans.  $\text{AlF}_6^{3-}$ ]
15. The shortest carbon-carbon bond distance is found in :  
(a) Diamond (b) Benzene (c) Ethane (d) Cyclopropane [Ans. (d)]
16.  $\text{HF}$ ,  $\text{H}_2\text{O}$ ,  $\text{HCl}$ ,  $\text{CCl}_4$  which is not liquid and why ?
17. Which is not see saw shaped ?  
(a)  $\text{SF}_4$  (b)  $\text{XeO}_2\text{F}_2$  (c)  $\text{XeOF}_2$  (d)  $\text{SiCl}_4$ . [Ans.  $\text{XeOF}_2$ ]

18.  $\text{H}_2^+ \text{H}_2^-$  have same bond order which is more stable ? [Ans.  $\text{H}_2^+$ ]
19.  $\text{B}_2$  has ten electrons but paramagnetic why ?
20. Draw the resonating structure of  $\bar{\text{NO}}_2$ .
21.  $\text{ClF}_3$  is T shaped but  $\text{BF}_3$  is planar. Explain.
22. Density of ice is less than water. Why ?
23. *o*-nitrophenol has less B.P. than *p*-nitrophenol. What is decreasing solubility in water ?
24.  $\text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+$   $\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+$ . Explain the type of bond.
25. Why do noble gas  $\text{Ne}_2$  does not exist but  $\text{Ne}_2^+$  exists ?
26.  $\text{H}_2\text{O}$  is liquid at room temperature but  $\text{H}_2\text{S}$  is gas why ?
27.  $\text{O}_2$  is paramagnetic but  $\text{O}_2^{2-}$  (peroxide ion) is diamagnetic why ?
28. What is the difference b/w  $\sigma$  (sigma) and  $\pi$  (pi) bond explain diagrammatically ?
29. (a)  $\text{F}_2$  and  $\text{Cl}_2$  which has less bond dissociation energy and why ?  
(b)  $\text{O}_2^+ \text{O}_2^-$  which is more stable and why ?
30. What is difference bond enthalpy and bond dissociation enthalpy ?

### 3-Mark Questions

1. Arrange in decreasing order of bond angle :
- (a)  $\text{H}_2\text{O}$ ,  $\text{H}_2\text{S}$ ,  $\text{H}_2\text{Se}$  [Ans.  $\text{H}_2\text{O} > \text{H}_2\text{S} > \text{H}_2\text{Se}$ ]
- (b)  $\text{NO}_2^+$ ,  $\text{NO}_2$ ,  $\text{NO}_2^-$  [Ans.  $\text{NO}_2^+ > \text{NO}_2 > \text{NO}_2^-$   
( $sp$ ) ( $sp^2$ ) ( $sp^2$ )]
- (c)  $\text{PF}_3$ ,  $\text{PH}_3$ ,  $\text{PCl}_3$  [Ans.  $\text{PCl}_3 > \text{PF}_3 > \text{PH}_3$ ]
2.  $\text{C}_2$  exists;  $\text{Be}_2$  does not why ? [Ans.  $\text{BO} \left[ \begin{array}{l} \text{Be}_2 = 0 \\ \text{C}_2 = 2 \end{array} \right]$ ]
3.  $\text{C}_2\text{H}_4$ ,  $\text{C}_2\text{H}_2$ ,  $\text{C}_2\text{H}_6$  arrange in decreasing :
- (a) Bond length C—C bond.
- (b) Bond energy of C—C bond.
4. (a)  $\text{C}_2\text{H}_2$  'C' is acidic.
- (b)  $\text{C}_2\text{H}_6$  'C' is electron donating.





- (d)  $\text{NO}_3^-$ ,  $\text{NO}_2^-$ , NO ('S' character of hybridization)
- (e)  $\text{BeCl}_2$ ,  $\text{BCl}_3$ ,  $\text{CCl}_4$ ,  $\text{PCl}_3$  (Bond angle)
2. Which has (a) zero dipole moment  $\text{CO}_2$ ;  $\text{SO}_2$ .  
 (b)  $sp$  hybridization  $\text{SO}_2$ ,  $\text{SO}_3$ , None.  
 (c) No octet rule  $\text{OF}_2$ ,  $\text{BCl}_3$ ,  $\text{H}_2\text{Se}$ .  
 (d) Zero bond order  $\text{Ne}_2$ ,  $\text{He}_2$  Both  
 (e) Paramagnetic character  $\text{Cl}^-$ , Br,  $\text{I}^-$
3. Which has both (a) Polar and nonpolar bond  $\text{H}_2\text{O}_2$   $\text{BCl}_3$ .  
 (b)  $sp$  and  $sp^3$  hybridization Propyne Ethyne  
 (c)  $\sigma$  and  $\pi$  bonds and  $sp$  hybridization  $\text{CH}_2 = \text{CH} = \text{CH}_2$   
 $\text{CH}_2 = \text{CH} - \text{CH}_3$   
 (d) Ionic and covalent bond, coordinate bonds as well  $\text{AgNO}_3$   $\text{AgF}$   $\text{Ag}_2\text{SO}_4$ .  
 (e) Coordinate and only  $\sigma$  bond  $\text{H}_3\text{O}^+$ ;  $\text{NH}_4^+$ ;  $\text{PH}_4^+$
4. (a) In  $dsp^2$  hybridization 'd' orbital is  $dx^2 - y^2$  or  $dz^2$   
 (b)  $sp^3d$  Hyb. 'd' orbital is  $dx^2 - y^2$  or  $dz^2$   
 (c)  $sp^3d^2$  Hyb.  $[\text{XeF}_4]$  Geometry is octahedral or square planar.  
 (d) CO has (a)  $sp$  (b)  $sp^2$  (c)  $sp$  with back bonding  
 (e)  $\text{PCl}_5$  has bond angles (i) Bond  $90^\circ$  and  $120^\circ$  (ii) Only  $90^\circ$   
 (iii) Only  $120^\circ$ .
5. Why (a)  $\text{BaSO}_4$  is insoluble although ionic in nature.  
 (b)  $\text{ClF}_3$  has only  $90^\circ$  bond angles.  
 (c)  $\text{SO}_2$  is angular but  $\text{SO}_3$  is planar.  
 (d)  $\text{NH}_3$ ,  $\text{PH}_3$  have same hybridization but different bond angle.  
 (e)  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$  loses  $4\text{H}_2\text{O}$  on heating but not fifth  $\text{H}_2\text{O}$ .
6. Explain the scheme of Hybridization in  $\text{C}_2\text{H}_4$ ,  $\text{C}_2\text{H}_6$ ,  $\text{C}_2\text{H}_2$ .
7.  $\text{CO}_3^{2-}$   $\text{NO}_3^-$   $\text{NO}_2^-$   $\text{SO}_4^{2-}$   $\text{PO}_4^{3-}$ . Draw their resonating structures and resonating hybrid structure? Why all bond angles in  $\text{CO}_3^{2-}$  are equal?
8. Draw the Lewis dot structure of (a)  $\text{Al}_2\text{O}_3$  (b)  $\text{Mg}_3\text{N}_2$  (c)  $\text{CCl}_4$  (d)  $\text{Na}_2\text{O}_2$

(e)  $\text{NCl}_3$ .

9. Draw the shapes (VSEPR) and geometries of :

$\text{SO}_3^{2-}$ ,  $\text{I}_3^-$ ,  $\text{XeO}_3$ ,  $\text{XeOF}_4$  and  $\text{NO}_2^-$ .

10. Arrange in properties as shown :

(a) HF, HCl, HBr, HI thermal stability

(b) LiF, LiCl, LiBr, LiI decreasing ionic character.

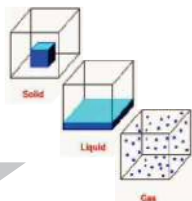
(c)  $\text{PH}_3$ ,  $\text{PCl}_3$  more covalent character.

(d)  $\text{O}_2 \rightarrow \text{O}_2^+$  bond length

$\text{N}_2 \rightarrow \text{N}_2^+$  bond length

In which case bond length decrease and why ?

(e) Why  $\text{CCl}_4$  has 4 dipole but net DM = zero.



## Chapter - 5

# States of Matter

### Part A

- On the basis of nature of intermolecular forces/molecular interactions, matter exists in three physical states : **solid, liquid and gas**.
- Intermolecular forces are the forces of attraction or repulsion between interacting particles (atoms and molecules). Attractive/repulsive intermolecular forces are known as van der Waal's forces.

Different types of van der Waal's forces are :

(a) **Dispersion forces or London forces** : The interaction which is present between two non polar molecules *i.e.*, between induced dipole and induces dipole is called dispersion forces, for example : noble gases.

(b) **Dipole-dipole forces** : The interaction which is present between molecules having permanent dipoles *i.e.*, between polar molecules, for example  $\text{NH}_3$ ,  $\text{HCl}$  etc.

(c) **Dipole-induced dipole forces** : The interaction which is present between a polar and non polar molecule.

- Boyle's law** : A constant temperature, the pressure of a fixed amount of gas, varies inversely with volume.

$$P \propto \frac{1}{V}$$

$$P_1 V_1 = P_2 V_2 \text{ [At constant temperature]}$$

- Charles's law** : At constant pressure, the volume of a fixed mass of gas is directly proportional to its absolute temperature.

$$V \propto T$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ [At constant pressure (p) and n]}$$

- Gay Lussac's law** : At constant volume, pressure of a fixed amount of

gas varies directly with the temperature.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad [\text{At constant volume (V) and } n]$$

- **STP (Standard Temperature and Pressure) :** STP means 273.15 K (0°C) temperature and 1 bar (*i.e.*, exactly  $10^5$  Pascal.) Volume occupied by 1 mole gas at STP = 22.7 L.

If pressure is taken in atm (atmosphere), then the standard molar volume is 22.4 L.

### Part-B

- **Ideal gas equation :**  $PV = nRT$

R is universal gas constant.

$$R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1} = 0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} = 0.083 \text{ L bar mol}^{-1} \text{ K}^{-1}$$

- **Combined gas law :**  $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

- **Density and molar mass of a gaseous substance :**

$$M = \frac{dRT}{P}$$

- **Dalton's law of partial pressure :** Dalton states that the total pressure exerted by the mixture of non reacting gases is equal to the sum of the partial pressures of individual of gases.

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots [\text{at constant T, V}]$$

$$P_1 = x_1 \times P_{\text{Total}} \quad \text{Here } x_1 \text{ is called mole fraction of the gas.}$$

$$P_{\text{dry gas}} = P_{\text{Total}} - \text{Aqueous tension.}$$

Here aqueous tension is the pressure exerted by water vapours.

- **Compressibility factor :** The extent of deviation of a real gas from an ideal behaviour is expressed in terms of compressibility factor,  $Z = \frac{PV}{nRT}$

For ideal gas,  $Z = 1$  at all temperatures and pressures. For real gases, greater is the deviation in the value of  $Z$  from 1, more is the deviation from ideal behaviour. When  $Z < 1$ , the gas is said to show negative deviation. This implies that gas is more compressible than expected from ideal behaviour. When  $Z > 1$ , the gas is said to show positive deviation and the gas is less compressible than expected from ideal behaviour.

At ordinary temperatures ( $T \geq 273 \text{ K}$ ), only  $\text{H}_2$  and  $\text{He}$  show positive deviations. However at low temperatures, even these gases show negative deviation *i.e.*,  $Z < 1$ . *For example*, in case of these gases, if  $T \ll 273 \text{ K}$ ,  $Z < 1$ .

- **Boyle temperature :** The temperature at which a real gas behaves like an ideal gas over an appreciable pressure range is called Boyle temperature or Boyle point.
- **Causes of deviation from ideal behaviour :** The following two assumptions of the kinetic theory of gases are faulty :
  - (a) The volume occupied by the gas molecules is negligible as compared to the total volume of the gas.
  - (b) The forces of attraction or repulsion between the gas molecules are negligible.

The above assumptions are correct only if the temperature is high or pressure is low.

**van der Waal's equation :**

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT \quad \text{for 1 mole of the gas}$$

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT \quad \text{for } n \text{ moles of the gas}$$

Here  $a$  and  $b$  are constants called van der Waal's constants.

- **Significance and units of van der Waal's constants :** ' $a$ ' gives the idea of the magnitude of attractive forces among the gas molecules. As correction in pressure is  $P = \frac{an^2}{V^2}$ , therefore  $a = (P \times V^2)/n^2 = \text{atm L}^2 \text{ mol}^{-2}$ .  
As correction in volume  $V = nb$ , therefore ' $b$ ' has the unit of  $\text{L mol}^{-1}$ . The near constancy in the volume of  $b$  shows that the gas molecules are incompressible.

### Part-C

- **Vapour pressure :** The pressure exerted by the vapours of a liquid, when it is in equilibrium with the liquid surface, at const. temperature.
- **Boiling temperature :** The temperature at which vapour pressure of a liquid is equal to the external pressure.
- At 1 atm, boiling temperature is called normal boiling point.
- At 1 bar boiling temperature is called standard boiling point.

- Vapour pressure of a pure liquid depends upon (i) intermolecular forces, (ii) Temperature.
- Surface tension is defined as force acting per unit length perpendicular to the line drawn on the surface. Its units is  $\text{Nm}^{-1}$ .
- **Effect of temperature on surface tension** : Surface tension decreases with increase in temperature with the increase in temperature, kinetic energy of molecules increases. As a result, intermolecular forces decreases and hence force acting per unit length decreases.
- **Viscosity** : It is defined as resistance offered to the flow of liquid due to internal friction between layers of fluids as they pass over each other.

$$F = \eta A \cdot \frac{du}{dx}$$

$\eta$  is called coefficient of viscosity.

- **Effect of temperature on viscosity** : Viscosity decrease with increase in temperature because with the increase in temperature the average kinetic energy increases and the intermolecular forces can be easily overcome.

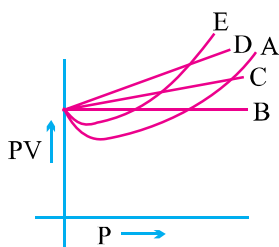
### 1-Mark Questions

#### Part-A

1. Name the intermolecular forces, which include the term. Vander Waal's forces. [Ans. LONDON forces, Dipole-Dipole, Dipole-induced dipole]
2. Define Boyle's law.
3. Write the condition in terms of temperature and pressure, under which all gases obey Charles's law. [Ans. High temperature, low pressure]
4. Mention the volume occupied by one mole of an ideal gas at STP.  
[Ans. 22.7 L]
5. Define the term absolute zero. [Ans. 0K]

#### Part-B

6. Define aqueous tension.
7. Mention the S.I. unit for the quantity  $\frac{PV^2T^2}{n^2}$ . [Ans.  $\text{Nm}^4 \text{K}^2 \text{mol}^{-1}$ ]
8. Define the term critical temperature.
9. Write any two postulates of kinetic molecular theory of gases.
10. Which curve in the figure represents the curve for ideal gas ? [Ans. B]



### Part-C

11. Define the term normal boiling point of a liquid.
12. Mention the factors on, which the vapour pressure of a pure liquid depends?
13. Define surface tension.
14. What do you understand by the term laminar flow ?
15. Out of Ethyl alcohol or Dimethylether, which one have higher vapour pressure at same temperature ?  
[Ans. Dimethylether]

### Part-A

#### 2 Marks Questions

1. What will be the minimum pressure required to compress  $500 \text{ dm}^3$  of air at 1 bar to  $200 \text{ dm}^3$  at  $30^\circ\text{C}$  ?  
[NCERT] [Ans. 2.5 bar]
2. Name the intermolecular force present in :  
(i)  $\text{H}_2\text{O}$  (ii)  $\text{HCl}$  [Ans. H-bonding, Dipole-Dipole]
3. Explain Avogadro's law.
4. Find the molar mass of a gas if 300 ml of this has mass of 0.368 g at STP.  
[Ans. 27.84 g/mol]
5. Why do real gases show deviation from ideal behaviour ? Write Van der Waal's equation for  $n$  moles of a gas.

### Part-B

6. Calculate the temperature of 4 mole of a gas occupying in  $5 \text{ dm}^3$  at 3.32 bar. ( $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$ )  
[NCERT] [Ans. 50K]
7. Calculate the volume occupied by 8.8 g of  $\text{CO}_2$  at  $31.1^\circ\text{C}$  and 1 bar pressure.  
[ $R = 0.83 \text{ bar L K}^{-1} \text{ mol}^{-1}$ ] [NCERT] [Ans. 5.05 L]
8. Explain the physical significance of vander Waal's parameter.

9. Compressibility factor 'Z' of a gas is given as  $Z = \frac{PV}{nRT}$ .
- (i) What is the value of Z for an ideal gas ?
  - (ii) For real gas, what will be the effect on value of Z above Boyle temperature? [Ans.  $Z = 1$ ,  $Z > 1$ ]
10. At 25°C and 760 mm Hg pressure a gas occupies 600 ml volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 mL. [NCERT] [Ans. 676.6 mm Hg]

### Part-C

11. Define the terms :
- (i) Standard boiling point.
  - (ii) Vapour pressure of a liquid.
12. Drops of liquid are spherical in nature. Explain. Mention the effect of temperature on surface tension.
13. Write the S.I. units of :
- (i) Surface tension.
  - (ii) Coefficient of viscosity.
14. Define viscosity. Mention the effect of temperature and pressure on viscosity of a liquid.
15. Explain :
- (i) Fire polishing of glass.
  - (ii) Liquid tend to rise in a capillary.

### Part-A

#### 3-Mark Questions

1. Define and explain Dalton's Law of partial pressure.
2. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 L volume, upto what volume can the ball on be expanded. [NCERT] [Ans. 11.35 L]
3. Calculate the total pressure in a mixture of 8g of dioxygen and 4g of dihydrogen confined in a vessel of 1dm<sup>3</sup> at 27°C.



[ $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$ ] [NCERT] [Ans. 56.025 bar]

4. 300 ml of oxygen gas at  $-10^\circ\text{C}$  are heated to  $10^\circ\text{C}$ . Find the volume of gas at  $10^\circ\text{C}$  if pressure remains constant. [Ans. 322.8 mL]
5. A gas at a pressure of 5 atm is heated from  $0^\circ$  to  $546^\circ\text{C}$  and is simultaneously compressed to one third of its original volume. Find the final pressure of the gas. [Ans. 45 atm]

### Part-B

6. Pressure of one gram of an ideal gas A at  $27^\circ\text{C}$  is found to be 2 bar. When 2g of another gas (ideal) B is introduced in the same flask at the same temperature the pressure becomes 3 bar. Find a relationship between their molecular masses. [Ans.  $M_B = 4M_A$ ]
7. What will be the pressure exerted by a mixture of 3.2g of methane and 4.4 g of carbon dioxide contained in a  $9 \text{ dm}^3$  flask at  $27^\circ\text{C}$ . [NCERT] [Ans. 0.82 atm]
8. A neon-dioxygen mixture contains 70.6 dioxygen and 167.5 neon. If the pressure of the mixture of gases in cylinder is 25 bar. What is the partial pressure of dioxygen and neon in the mixture. [Ans. 5.25 bar, 17.75 bar]
9. With the help of a gas laws, deduce an expression for the ideal gas equation. What is the utility of the gas equation ?
10. A vessel of 120 mL capacity contains a certain mass of a gas at  $20^\circ\text{C}$  and 750 mm pressure the gas was transferred to a vessel whose volume is 180 mL. Calculate the pressure of the gas at  $20^\circ\text{C}$ . [Ans. 500 mm]

### Part-C

11. Explain :
- (i) Liquid at higher altitudes boil at low temperature.
  - (ii) In hospital surgical instruments are sterilised in autoclaves.
  - (iii) Out of, alcohols and ethers of comparable mass which one have higher boiling points ?
12. (i) Define surface energy in relation to surface tension.
- (ii) Name the temperature at which the density of water is maximum.
  - (iii) Moist soil grains are pulled together. Explain.
13. (i) Define the term coefficient of viscosity. Name the unit of viscosity coefficient in cgs system.
- (ii) Give the difference between boiling and evaporation.

14. Explain :

- (i) Tea or coffee is sipped from the saucer, when it is quite hot.
- (ii) Liquids possesses fluidity.

15. Which among the following will have ?

- (i) HCl or H<sub>2</sub>O (Higher boiling point)
- (ii) Ether or water (Higher viscosity)
- (iii) Br<sub>2</sub>(l) or water (Lower surface tension)

### 5-Mark Questions

1. Mention the intermolecular forces present between :

- (a) H<sub>2</sub>O and alcohol (b) Cl<sub>2</sub> and CCl<sub>4</sub> (c) He and He atoms (d) Na<sup>+</sup> ion and H<sub>2</sub>O (e) HBr and HBr.

2. (a) Find the pressure of 4g of O<sub>2</sub> and 2g of H<sub>2</sub> confined in a bulb of 1 litre at 0°C. [Ans. 25.215 atm]

(b) What is the molar volume of a gas at SATP conditions ?

(c) Define and explain Gay Lussac's law.

### Part-B

3. (a) For Dalton's law of pressure derive the expression  $P_{\text{gas}} = X_{\text{gas}} P_{\text{total}}$ .  
(b) A 2-L flask contains 1.6 g of methane and 0.5 g of hydrogen at 27°C. Calculate the partial pressure of each gas in the mixture and hence, calculate the total pressure.

[Ans.  $p_{\text{CH}_4} = 1.23 \text{ atm}$ ,  $p_{\text{H}_2} = 3.079 \text{ atm}$ ,  $P_{\text{total}} = 4.31 \text{ atm}$ .]

4. (a) Using Vander waal's equation calculate the constant 'a' when two moles of a gas confined in a four litre flask exerts a pressure of 11.0 atm. at a temperature of 300 K. The value of 'b' is 0.05 litre mol<sup>-1</sup>.

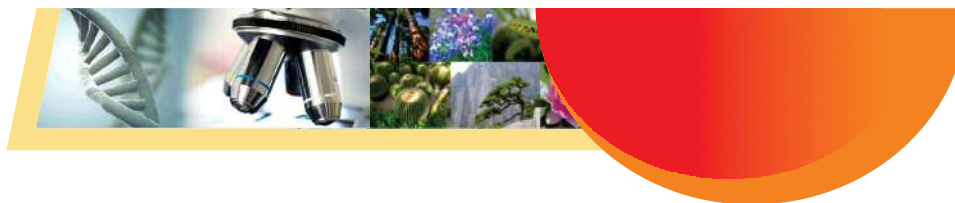
[Ans. 6.49 atm L<sup>2</sup> mol<sup>-2</sup>]

5. (a) Mention the factors on which vapour pressure of a pure liquid depends.  
(b) Define boiling point of a liquid.  
(c) Which will have higher viscosity Glycerol or Ethylene glycol.

[Ans. Glycerol]

- (d) Surface tension of a liquid ..... with increase in the magnitude of intermolecular forces.

[Ans. Increases]



## Chapter - 6

# Thermodynamic

- **System** : Specific part of universe in which observations are made.
- **Surroundings** : Everything which surrounds the system.
- **Types of the System** :
  - (i) **Open System** : Exchange both matter and energy with the surroundings. For example : Reactant in an open test tube.
  - (ii) **Closed System** : Exchange energy but no matter with the surroundings. For example : Reactants in a closed vessel.
  - (iii) **Isolated System** : Neither exchange energy nor matter with the surroundings. For example : Reactants in a thermos flask.
- Please note no system is perfectly isolated.
- **Thermodynamic Processes** :
  - (i) Isothermal process :  $\Delta T = 0$
  - (ii) Adiabatic process :  $\Delta q = 0$
  - (iii) Isobaric process :  $\Delta P = 0$
  - (iv) Isochoric process :  $\Delta V = 0$
  - (v) Cyclic process :  $\Delta U = 0$
  - (vi) Reversible process : Process which proceeds infinitely slowly by a series of equilibrium steps.
  - (vii) Irreversible process : Process which proceeds rapidly and the system does not have chance to achieve equilibrium.
- **Extensive Properties** : Properties which depend upon the quantity or size of matter present in the system. For example : mass, volume, internal energy, enthalpy, heat capacity, work etc.

- **Intensive Properties :** Properties which do not depend upon the quantity or size of matter present in the system. For example : temperature, density, pressure, surface tension, viscosity, refractive index, boiling point, melting point etc.
- **State Functions :** The variables of functions whose value depend only on the state of a system or they are path independent. For example : pressure (P), volume (V), temperature (T), enthalpy (H), free energy (G), internal energy (U), entropy (S), amount (n) etc.
- **Internal Energy :** It is the sum of all kind of energies possessed by the system.
- **First Law of Thermodynamics :** “The energy of an isolated system is constant.”

$$\text{Mathematical Form : } \Delta U = q + w$$

- **Sign Conventions for Heat (q) and Work (w) :**

- (i)  $W = +ve$ , if work is done on system
- (ii)  $W = -ve$ , if work is done by system
- (iii)  $q = +ve$ , if heat is absorbed by the system
- (iv)  $q = -ve$ , if heat is evolved by the system

- Work of Expansion/compression :  $w = -P_{ext}(V_f - V_i)$

- **Work done in Isothermal Reversible Expansion of an Ideal Gas :**

$$w_{rev} = -2.303 nRT \log \frac{V_f}{V_i}$$

$$\text{Or, } w_{rev} = -2.303 nRT \log \frac{P_i}{P_f}$$

- Significance of  $\Delta H$  and  $\Delta U$  :  $\Delta H = q_p$  and  $\Delta U = q_v$
- **Relation** between  $\Delta H$  and  $\Delta U$ :  $\Delta H = \Delta U + (n_p - n_r)RT$  for gaseous reaction.
  - (i)  $\Delta H = \Delta U$  if  $(n_p - n_r)$  is zero; e.g.,  $H_2(g) + I_2(g) \rightarrow 2HI(g)$
  - (ii)  $\Delta H > \Delta U$  if  $(n_p - n_r)$  is positive; e.g.,  $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$
  - (iii)  $\Delta H < \Delta U$  if  $(n_p - n_r)$  is negative; e.g.,  $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
- **Heat capacity (C) :** Amount of heat required to raise the temperature of a substance by  $1^\circ\text{C}$  to 1 K.

$$q = C\Delta T$$

- **Specific heat capacity ( $C_s$ ) :** Amount of heat required to raise the

temperature of 1g of a substance by 1°C or 1K.

$$q = C_s \times m \times \Delta T$$

- **Molar Heat Capacity ( $C_m$ )** : Amount of heat required to raise the temperature of 1 mole of a substance by 1°C or 1K.

$$q = C_m \times n \times \Delta T$$

- **Standard State of a Substance** : The standard state of a substance at a specified temperature is its, pure form at 1 bar.
- **Standard Enthalpy of Formation ( $\Delta_f H^\circ$ )** : Enthalpy change accompanying the formation of one mole of a substance from its constituent elements under standard condition of temperature (normally 298 K) and pressure (1 bar).
  - $\Delta_f H^\circ$  of an element in standard state is taken as zero.
  - Compounds with – ve value of  $\Delta_f H^\circ$  are more stable than their constituents.
  - $\Delta_r H^\circ = \sum_i a_i \Delta_f H^\circ (\text{products}) - \sum_i b_i \Delta_f H^\circ (\text{reactants})$  : Where ‘a’ and ‘b’ are coefficients of products and reactants in balanced equation.
- **Standard Enthalpy of Combustion ( $\Delta_c H^\circ$ )** : Enthalpy change accompanying the complete combustion of one mole of a substance under standard conditions (298 K, 1 bar)
- **Hess’s Law of Constant Heat Summation** : The total enthalpy change of a reaction remains same whether it takes place in one step or in several steps.
- **Bond Dissociation Enthalpy** : Enthalpy change when one mole of a gaseous covalent bond is broken to form products in gas phase. For example:  $\text{Cl}_2(\text{g}) \rightarrow 2\text{Cl}(\text{g})$ ;  $\Delta_{\text{Cl-Cl}} H^\circ = 242 \text{ kJ/mol}^{-1}$ .
- For diatomic gaseous molecules; Bond enthalpy = Bond dissociation Enthalpy = Atomization Enthalpy
- For Polyatomic gaseous molecules; Bond Enthalpy = Average of the bond dissociation enthalpies of the bonds of the same type.
- $\Delta_r H^\circ = \sum \Delta_{\text{bond}} H^\circ (\text{Reactants}) - \sum \Delta_{\text{bond}} H^\circ (\text{Products})$
- **Spontaneous Reaction** : A reaction which can take place either of its own or under some initiation.

- **Entropy (S)** : It is measure of degree of randomness or disorder of a system.
- $\Delta S_{sys} = \frac{(q_{rev})_{sys}}{\Delta T} = \frac{(\Delta H)_{sys}}{\Delta T}$
- Unit of Entropy =  $\text{JK}^{-1} \text{mol}^{-1}$
- **Second Law of Thermodynamics** : For all the spontaneous processes totally entropy change must positive.  

$$\Delta S_{total} = \Delta S_{sys} + \Delta S_{surr} > 0$$
- **Gibbs Helmholtz Equation for determination of Spontaneity** :  

$$\Delta G = \Delta H - T\Delta S$$
  - (i) If  $\Delta G = -ve$ , the process is spontaneous
  - (ii) If  $\Delta G = +ve$ , the process is nonspontaneous
  - (iii) If  $\Delta G = 0$ , the process is in equilibrium
- **Relation between Gibbs Energy Change and Equilibrium Constant** :  

$$\Delta G^0 = -2.303 RT \log K_c$$
- **Third law of thermodynamic** : The entropy of a perfectly crystalline solid at absolute zero (0 K) is taken to be zero.

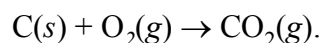
### 1-Mark Questions

#### Thermodynamic Concepts and First Law of Thermodynamics

1. Day temp. Name the thermodynamic system to which following belong:  
 (i) Human body (ii) Milk in Thermos flask (iii) Tea in steel kettle
2. Identity State functions out of the following : Enthalpy, Entropy, Heat, Temperature, Work, Free energy.
3. Give two examples of state functions.
4. Write the mathematical statement of first law of thermodynamics.

#### Work, Enthalpy, Internal Energy and Heat Capacity

5. Predict the internal energy change for an isolated system ? [**Ans.** Zero]
6. Why  $\Delta H$  is more significant than  $\Delta U$  ?
7. Write one example each of extensive and intensive properties.
8. Write a chemical equation in which  $\Delta H$  and  $\Delta U$  are equal.
9. Write the relationship between  $\Delta H$  and  $\Delta U$  for the reaction :



### Different types of Enthalpy Changes and Hess's Law

10. Define standard enthalpy of formation.
11. Why is the standard enthalpy of formation of diamond not zero although it is an element ?
12. The enthalpy of atomization of  $\text{CH}_4$  is  $1665 \text{ kJ mol}^{-1}$ . What is the bond enthalpy of C-H bond ? [Ans. 416.25 kJ]
13. Identify the species for which  $\Delta_f H^\circ = 0$ , at 298 K :  $-\text{Br}_2$ ,  $\text{Cl}_2$ ,  $\text{CH}_4$ .  
[Hint :  $\text{Cl}_2$  ( $\text{Br}_2$  is liquid at 298K)]

### Spontaneity, Entropy, Second and Third Law of Thermodynamics

14. For the reaction  $2\text{Cl(g)} \rightarrow \text{Cl}_2\text{(g)}$ ; what are the sign of  $\Delta H$  and  $\Delta S$  ?
15. For an isolated system  $\Delta U = 0$ , what will be  $\Delta S$  ?
16. Why entropy of steam is more than that of water at its boiling point ?
17. Out of Diamond and Graphite which has higher entropy ?
18. Write an example of endothermic spontaneous reaction.
19. State second law of thermodynamics.
20. State third law of thermodynamics.
21. Which has more entropy ? 1 mol  $\text{H}_2\text{O(l)}$  at  $25^\circ\text{C}$  or 1 mol  $\text{H}_2\text{O(l)}$  at  $35^\circ\text{C}$ .
22. At what temperature the entropy of a perfectly crystalline solid is zero?

### Gibbs Energy and Spontaneity, Effect of Temperature on Spontaneity

23. For a certain reaction  $\Delta G^\circ = 0$ , what is the value of  $K_c$  ?
24. How can a non spontaneous reaction be made spontaneous ?
25. For a reaction both  $\Delta H$  and  $\Delta S$  are negative. Under what conditions does the reaction occur.

### 2 Marks Question

### Thermodynamic Concepts and First Law of Thermodynamics

1. In a process 701 J of heat is absorbed by a system and 394 J work is done by the system. What is the change in internal energy for the process ?  
[Ans. 307 J]

2. Neither  $q$  nor  $w$  is state functions but  $q + w$  is a state function. Explain.

### Work, Enthalpy, Internal Energy and Heat Capacity

3. Classify the following as extensive or intensive properties :  
Heat capacity, Density, Temperature, Molar heat capacity.
4. Derive the relationship between  $\Delta H$  and  $\Delta U$ .
5. Derive the relationship  $C_p - C_v = R$ .
6. A 1.25g sample of octane ( $C_{18}H_{18}$ ) is burnt in excess of oxygen in a bomb calorimeter. The temperature of the calorimeter rises from 294.05 to 300.78K. If heat capacity of the calorimeter is  $8.93 \text{ kJ K}^{-1}$ . Find the heat transferred to calorimeter. [Ans. 0.075 kJ]
7. Show that for an ideal gas, the molar heat capacity under constant volume conditions is equal to  $3/2 R$ .
8. Expansion of a gas in vacuum is called free expansion. Calculate the work done and change in internal energy when 1 mol of an ideal gas expands isothermally from 1 L to 5 L into vacuum.

### Different types of Enthalpy Changes and Hess's Law

9. State and explain Hess's Law of Constant Heat Summation with a suitable example.
10. Derive the relationship between  $\Delta H$  and  $\Delta U$ . Given,  $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ ;  $\Delta_f H^\circ = -92.4 \text{ kJ mol}^{-1}$ ; What is the standard enthalpy of formation of  $NH_3$  gas ? [Ans.  $-46.2 \text{ kJ mol}^{-1}$ ]
11. Calculate the enthalpy change for the reaction :  $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$ . Given the bond enthalpies  $H_2$ ,  $Br_2$  and  $HBr$  are  $435 \text{ kJ mol}^{-1}$ ,  $192 \text{ kJ mol}^{-1}$  and  $368 \text{ kJ mol}^{-1}$  respectively. [Ans.  $-109 \text{ kJ mol}^{-1}$ ]
12. Is the bond dissociation enthalpy of all the four C-H bonds in  $CH_4$  same? Give reason in support of your.

### Spontaneity, Entropy, Second and Third Law of Thermodynamics

13. Define the term entropy. Write its unit. How does entropy of a system change on increasing temperature?
14. Dissolution of ammonium chloride in water is endothermic but still it dissolves in water readily. Why ?
15. Calculate the entropy change in the surroundings when 1.00 mol of  $HI(OI)$  is formed under standard conditions;  $\Delta_f H^\circ = -286 \text{ kJ mol}^{-1}$ .



[Ans. 959.7 J K<sup>-1</sup> mol<sup>-1</sup>]

16. The enthalpy of vaporization of a liquid is 30 kJ mol<sup>-1</sup> and entropy of vaporization is 75 J K<sup>-1</sup> mol<sup>-1</sup>. Calculate the boiling point of liquid at 1 atm. [Ans. 400 K]

### Gibbs Energy and Spontaneity, Effect of Temperature on Spontaneity

17. The equilibrium constant for a reaction is 10. What will be the value of  $\Delta G^\circ$ ?  $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$ ,  $T = 300 \text{ K}$ . [Ans. - 5.527 kJ mol<sup>-1</sup>]
18. Derive the relationship,  $\Delta G = -T\Delta S_{\text{total}}$  for a system.
19. The  $\Delta H$  and  $\Delta S$  for  $2\text{Ag}_2\text{O}(s) \rightarrow 4\text{Ag}(s) + \text{O}(g)$  are given + 61.17 kJ mol<sup>-1</sup> and + 132 K<sup>-1</sup> mol<sup>-1</sup> respectively. Above what temperature will the reaction be spontaneous? [Ans. > 463.4 K]

### 3 Mark Questions

### Thermodynamic Concepts and First Law of Thermodynamics

1. Differentiate between the following (with examples) :
- (i) Open and Closed System.
  - (ii) Adiabatic and Isothermal process
  - (iii) State function and path function

### Work, Enthalpy, Internal Energy and Heat Capacity

2. Calculate the maximum work obtained when 0.75 mole of an ideal gas expands isothermally and reversibly at 27°C from a volume of 15 L to 25 L. [Ans. - 955.7 J]
3. Calculate the number of kJ necessary to raise the temperature of 60 g of aluminium from 35 to 55°C. Molar heat capacity of Al is 24 J mol<sup>-1</sup> K<sup>-1</sup>. [Ans. 1.067 kJ]
4. The reaction of cyanamide,  $\text{NH}_2\text{CN}(s)$ , with Dioxygen was carried out in a bomb calorimeter, and  $\Delta U$  was found to be - 742.7 kJ mol<sup>-1</sup> at 298K. Calculate Enthalpy change for the reaction at 298K,  $\text{NH}_2\text{CN}(s) + \frac{3}{2} \text{O}_2(g) \rightarrow \text{N}_2(g) + \text{CO}_2(g) + \text{H}_2\text{O}(l)$  [Ans. - 741.5 kJ mol<sup>-1</sup>]

### Different types of Enthalpy Changes and Hess's Law

5. The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are - 890.3 kJ mol<sup>-1</sup>, - 393.5 kJ mol<sup>-1</sup> and - 285.8 kJ mol<sup>-1</sup> respectively.

Calculate enthalpy of formation of methane gas. [**Ans.**  $-74.8 \text{ kJ mol}^{-1}$ ]

6. Explain the Born Haber Cycle to determine the lattice enthalpy of NaCl.
7. Enthalpies of formation of  $\text{CO(g)}$ ,  $\text{CO}_2\text{(g)}$ ,  $\text{N}_2\text{O(g)}$  and  $\text{N}_2\text{O}_4\text{(g)}$  are  $-110$ ,  $-393$ ,  $81$  and  $9.7 \text{ kJ mol}^{-1}$  respectively. Find the value of  $\Delta_r H^\circ$  for the reaction;  $\text{N}_2\text{O}_4\text{(g)} + 3\text{CO(g)} \rightarrow \text{N}_2\text{O(g)} + 3\text{CO}_2\text{(g)}$

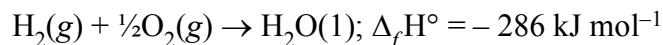
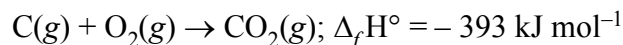
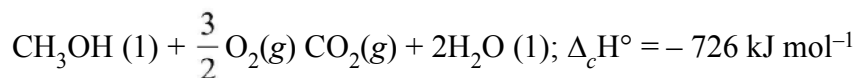
[**Ans.**  $-777.7 \text{ kJ mol}^{-1}$ ]

8. The combustion of 1 mol of benzene takes place at  $298\text{K}$ . After combustion  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are formed and  $3267 \text{ kJ mol}^{-1}$  of heat is liberated. Calculate  $\Delta_f H^\circ(\text{C}_6\text{H}_6)$ .

Given :  $\Delta_f H^\circ(\text{CO}_2) = -286 \text{ kJ mol}^{-1}$ ,  $\Delta_f H^\circ(\text{H}_2\text{O}) = -393 \text{ kJ mol}^{-1}$

[**Ans.**  $48.51 \text{ kJ mol}^{-1}$ ]

9. Calculate the standard enthalpy of formation of  $\text{CH}_3\text{OH (l)}$  from the following data :



[**Ans.**  $-239 \text{ kJ mol}^{-1}$ ]

### Spontaneity, Entropy, Second and Third Law of Thermodynamics

10. For oxidation of iron,  $4\text{Fe(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Fe}_2\text{O}_3$  entropy change is  $-549.4 \text{ J K}^{-1} \text{ mol}^{-1}$  at  $298 \text{ K}$ . In spite of negative entropy change of this reaction, why is the reaction spontaneous ? ( $\Delta_r H^\circ$  for this reaction is  $-1648 \text{ kJ mol}^{-1}$ )

[**Ans.**  $\Delta S_{\text{total}} = +4980.6 \text{ J K}^{-1} \text{ mol}^{-1}$ ]

11. Give reasons :

- (i) Evaporation of water is an endothermic process but it is spontaneous,
- (ii) A real crystal has more entropy than an ideal crystal,
- (iii) Entropy of universe is increasing.

### Gibbs Energy and Spontaneity, Effect of Temperature on Spontaneity

12. For the reaction at  $298 \text{ K}$ ,  $2\text{A} + \text{B} \rightarrow \text{C}$ ;  $\Delta H = 400 \text{ kJ mol}^{-1}$ ,  $\Delta S = 0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}$ . At what temperature will the reaction become spontaneous considering  $\Delta H$  and  $\Delta S$  to be constant over the temperature range.

[**Ans.**  $T > 2000 \text{ K}$ ]

13. Reaction  $X \rightarrow Y$ ;  $\Delta H = +ve$  is spontaneous at temperature “T”. Determine
- Sign of  $\Delta S$  for this reaction,
  - Sign of  $\Delta G$  for  $Y \rightarrow X$
  - Sign of  $\Delta G$  at a temperature  $< T$ .

### 5-Mark Questions

#### Thermodynamic Concepts and First Law of Thermodynamics

- What is reversible process in Thermodynamics?
  - Name the thermodynamic processes for which : (i)  $q = 0$  (ii)  $\Delta U = 0$  (iii)  $\Delta V = 0$  (iv)  $\Delta P = 0$ .
  - Water decomposes by absorbing 286.2 kJ of electrical energy per mole. When  $H_2$  and  $O_2$  combine to form one mole of  $H_2O$ , 286.2 kJ of heat is produced. Which thermodynamic law is proved? Write its statement.

#### Work, Enthalpy, Internal Energy and Heat Capacity .

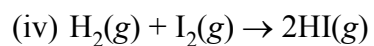
- Although heat is a path function but heat absorbed by the system under certain specific conditions is independent of path. What are those conditions ? Explain. [Hint :  $q_v = \Delta U$  and  $q_p = \Delta H$ ]
  - It has been found that 221.4 J is needed to heat 30g of ethanol from  $15^\circ C$  to  $18^\circ C$ . Calculate (a) specific heat capacity, and (b) molar heat capacity of ethanol. [Ans. (a)  $2.46 J g^{-1} ^\circ C^{-1}$ , (b)  $113.2 J mol^{-1} ^\circ C^{-1}$ ]

#### Different types of Enthalpy Changes and Hess's Law

- Differentiate the terms Bond dissociation enthalpy and Bond Enthalpy.
  - Calculate enthalpy change for the process  $CCl_4(g) \rightarrow C(g) + 4Cl(g)$  and calculate Bond enthalpy of C-Cl bond in  $CCl_4$ . Given :  $\Delta_{vap} H^\circ(CCl_4) = 30.5 kJ mol^{-1}$ ;  $\Delta_f H^\circ(CCl_4) = -135.5 kJ mol^{-1}$ ;  $\Delta_a H^\circ(C) = 715 kJ mol^{-1}$  and  $\Delta_a H^\circ(Cl_2) = 242 kJ mol^{-1}$ . [Ans.  $1304 kJ mol^{-1}$ ,  $326 kJ mol^{-1}$ ]

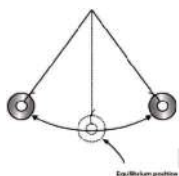
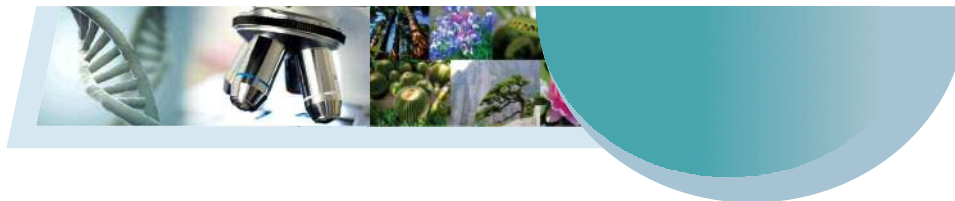
#### Spontaneity, Entropy, Second and Third Law

- Predict the sign of  $\Delta S$  for the following changes :
    - Freezing of water,
    - $C(\text{graphite}) \rightarrow C(\text{diamond})$
    - $H_2(g)$  at 298 k and 1 bar  $\rightarrow H_2(g)$  at 298 k and 10 bar



### Gibbs Energy and Spontaneity, Effect of Temperature of Spontaneity

5. (i) Define Gibbs energy. Give its mathematical expression. What is Gibbs's energy criteria of spontaneity.
- (ii) For the reaction :  $2\text{A}(\text{g}) + \text{B}(\text{g}) \rightarrow 2\text{D}(\text{g})$ ,  $\Delta U^\circ = -10.5 \text{ kJ}$  and  $\Delta S^\circ = -44.1 \text{ J K}^{-1}$ . Calculate  $\Delta_r G^\circ$  for the reaction, and predict whether it will occur spontaneously. [Ans.  $\Delta_r G^\circ = +0.16 \text{ kJ}$ , Non spontaneous]

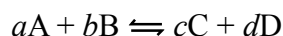


## Chapter - 7

# Equilibrium

- **Equilibrium** : It is a state in a process when two opposing processes (forward and reverse) occur simultaneously at the same rate.  
The free energy change at equilibrium state is zero *i.e.*,  $\Delta G = 0$ .

- **Equilibrium constant** : For a general reaction :



$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b} \text{ and } K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

- **Relationship between  $K_p$  and  $K_c$  :**

$$K_p = K_c (RT)^{\Delta n_g}$$

$$\Delta n_g = n_p(g) - n_r(g)$$

Magnitude of equilibrium constant depends upon the way in which a reaction is written :

Chemical equation	Equilibrium constant
$aA + bB \rightleftharpoons cC + dD$	$K$
$cC + dD \rightleftharpoons aA + bB$	$K_1 = \frac{1}{K}$
$naA + nbB \rightleftharpoons ncC + ndD$	$K_2 = K^n$
$\frac{1}{n}aA + \frac{1}{n}bB \rightleftharpoons \frac{1}{n}cC + \frac{1}{n}dD$	$K_3 = K^{1/n}$

- **Predicting the direction of reaction :**

If  $Q_c = K_c \Rightarrow$  the reaction is in a state of equilibrium.

$Q_c > K_c \Rightarrow$  the reaction proceeds in reverse direction.

$Q_c < K_c \Rightarrow$  the reaction proceeds in forward direction.

- **Ostwald's dilution law** : Degree of dissociation of weak electrolyte,

$$\alpha = \sqrt{\frac{K}{C}}$$

- **Ionic Product of water ( $K_w$ )** =  $[H_3O^+][OH^-] = 10^{-14}$  at 298K

**Le-Chatelier's Principle** : When a system of equilibrium is subjected to a change in temperature, pressure or concentration, the equilibrium shifts itself in such a way so as to undo or counterout the effect of change.

- **Outcomes of Le-Chatelier's Principle**

Change at equilibrium	Shift in equilibrium
Increase in temperature	Endothermic direction
Decrease in temperature	Exothermic direction
Increase in pressure	Towards lesser gaseous moles
Decrease in pressure	Towards greater gaseous moles
Increase in Conc. of reactants	Forward direction
Increase in Conc. of products	Reverse direction

- **Conjugate Acid or Base** :

Species —  $H^+$  = Conjugate base

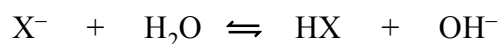
Species +  $H^+$  = Conjugate acid

- **pH of solution** :

$$pH = -\log [H_3O^+] \text{ or } [H^+] = 10^{-pH}$$

$$pH + pOH = pK_w = 14 \text{ at } 298K$$

- **Common ion effect** : The depression of ionisation of weak electrolyte by the presence of common ion from a strong electrolyte is called common ion effect. For example degree of dissociation of  $NH_4OH$  decreases in the presence of strong electrolyte  $NH_4Cl$ .
- **Hydrolysis of salts and pH of their solutions** : Hydrolysis of salt is defined as the reaction of cation or anion with water as a result of which the pH of water changes.
  1. Salts of strong and strong bases (*e.g.*,  $NaCl$ ) do not hydrolyse. The solution pH will be 7.
  2. Salts of weak acids and strong bases (*e.g.*,  $CH_3COONa$ ) hydrolyse,  $pH > 7$  (The anion acts as a base).

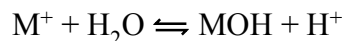


(Weak acid)

(Strong base)

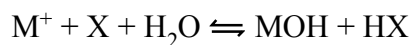
$$\text{pH} = 7 + \frac{1}{2} (\text{pK}_a + \log C)$$

3. Salt of strong acids and weak bases (*e.g.*,  $\text{NH}_4\text{Cl}$ ) hydrolyse,  $\text{pH} < 7$ .  
(The cation acts as an acid).



$$\text{pH} = 7 - \frac{1}{2} (\text{pK}_a + \log C)$$

4. Salt of weak acids and weak base (*e.g.*,  $\text{CH}_3\text{COONH}_4$ ) hydrolyse. The cation acts as an acid and anion as a base but whether the solution is acidic or basic depends upon the relative values of  $\text{K}_a$  and  $\text{K}_b$  for these ions.



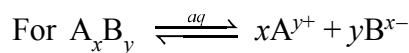
$$\text{pH} = 7 + \frac{1}{2} (\text{pK}_a - \text{pK}_b)$$

- **Buffer solutions** : The solutions, which resist the change in pH on dilution or addition of small amounts of acid or base, are called buffer solutions.
- **Basic buffer** : Solution of weak base and its salt with strong acid, For *e.g.*,  $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$
- **Acidic buffer** : Solution of weak acid and its salt with strong base, For *e.g.*,  $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$ .
- **Henderson's Equation for the pH of Buffer solution—**

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Salt}]}{[\text{Acid}]} \quad (\text{for acidic buffer})$$

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{Salt}]}{[\text{Base}]} \quad (\text{for basic buffer})$$

- **Solubility Product ( $\text{K}_{sp}$ )** : The equilibrium constant that represent the equilibrium between undissolved salt (solute) and its ions in a saturated solution is called solubility product constant ( $\text{K}_{sp}$ ).



$$\text{K}_{sp} = [\text{A}^{y+}]^x [\text{B}^{x-}]^y$$

If ionic product  $< \text{K}_{sp}$  ; salt remain dissolve.

If ionic product  $> \text{K}_{sp}$  ; salt will be precipitated.

- **Relationship between solubility (S) and solubility product ( $\text{K}_{sp}$ ).**

$$K_{sp} = x^x \cdot y^y \cdot S^{x+y}$$

For binary salts (*e.g.*, AgCl, AgBr, AgI)

$$K_{sp} = S^2$$

For Ternary salts (*e.g.*,  $PbI_2$ )

$$K_{sp} = 4S^3$$

## Chemical Equilibrium

### 1-Mark Questions

1. Define physical equilibrium. Give an example also.
2. Why is there a fizz when soda water bottle is opened ?
3. Justify the statement : ‘Both physical and chemical equilibria are dynamic in nature’
4. State Henry’s law.
5. In a reversible reaction, the two substances are in equilibrium. If the concentration of each one is reduced to half, then what is the effect on the equilibrium constant ?
6.  $K_1$  and  $K_2$  are equilibrium constant for reactions (1) and (2)  

$$N_2(g) + O_2(g) \rightleftharpoons 2 NO(g)$$

$$NO(g) \rightleftharpoons 1/2 N_2(g) + 1/2 O_2(g)$$
 then, calculate the relation between  $K_1$  and  $K_2$
7. Write the equilibrium constant expression for the following reaction :  

$$3 Fe(s) + 4 H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4 H_2(g)$$
8. Classify the equilibrium as homogenous or heterogenous :  

$$CH_3COOC_2H_5(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + C_2H_5OH(aq)$$
9. 
$$K_p = \frac{(P_{NH_3})}{(P_{N_2})^{1/2} (P_{H_2})^{3/2}}$$
 Write the balanced chemical equation corresponding to the above expression.
10. Give the direction in which the reaction would proceed if  $Q_c > K_c$
11.  $Hb(s) + O_2(g) \rightleftharpoons HbO_2(s)$   
 Predict the direction in which equilibrium gets shifted if partial pressure of  $O_2(g)$  is lowered.
12. Discuss the position of equilibrium if the following reaction is carried out in the presence of catalyst.  

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$



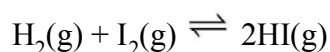
13. In which case the yield of the product will be more.



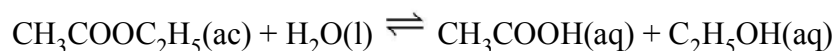
14. Value of  $K_c$  at 500 K for  $A + 2B \rightleftharpoons C$  is 10.6

Determine the value of  $K_c$  at the same temperature for  $C \rightleftharpoons A + 2B$

15. Why do we sweat more on a humid day ?  
16. Why does a catalyst not affect the magnitude of equilibrium constant ?  
17. Write the relation between  $K_p$  and  $K_c$  for the given reaction.



18. Write the expression of  $K_c$  for the following reaction



19. Write the expression of  $K_p$  for the following reaction



20. For an exothermic reaction, what happens to the equilibrium constant if temperature is increase ?  
21.  $K_a$  for  $HA_1$  is 10 and  $K_a$  for  $HA_2$  is 12  
Which acid is stronger.

22. For tribasic acid  $K_{a_1} > K_{a_2} > K_{a_3}$

What will happen to the acid strength of polyprotic acid if protons are lost ?

23.  $C(s) \rightleftharpoons C(s)$

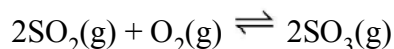
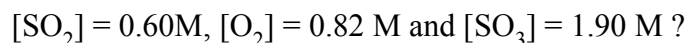
Diamond (Density =  $3.5 \text{ g/cm}^3$ ) Graphite (Density =  $2.3 \text{ g/cm}^3$ )

What will be the effect of increasing the pressure in this equilibrium.

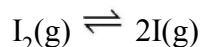
24. Under what condition does the active mass becomes equal to the molarity ?  
25. In which expression of  $K_c$  or  $K_{eq}$ , that only solutions and gases are involved ?  
26. What will be the change in the direction of equilibrium if Ne gas is added in a reaction at constant volume.  
27. What will be the effect on the boiling point of liquid if pressure is increased ?

## 2-Mark Questions

1. What is  $K_c$  for the following equilibrium when the equilibrium concentration of each substance is :



2. At a certain temperature and total pressure of  $10^5\text{ Pa}$ , iodine vapour contains 40% by volume of I atoms



Calculate  $K_p$  for the equilibrium. [NCERT] [ $2.67 \times 10^4\text{ Pa}$ ]

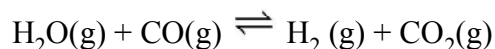
3. Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression ?

4. A sample of  $\text{HI}(\text{g})$  is placed in flask at a pressure of 0.2 atm. At equilibrium the partial pressure of  $\text{HI}(\text{g})$  is 0.04 atm.

What is  $K_p$  for the given equilibrium ?

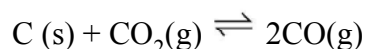


5. One mole of  $\text{H}_2\text{O}$  and one mole of  $\text{CO}$  are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water (by mass) reacts with  $\text{CO}$  according to the equation,



Calculate the equilibrium constant for the reaction. [NCERT] [0.444]

6. At 1127 K and 1 atm pressure, a gaseous mixture of  $\text{CO}$  and  $\text{CO}_2$  in equilibrium with solid carbon has 90.55%  $\text{CO}$  by mass



Calculate  $K_c$  for this reaction at the above temperature. [NCERT] [0.153]

7. Describe the effect of :

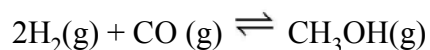
(a) addition of  $\text{H}_2$

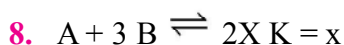
(b) addition of  $\text{CH}_3\text{OH}$

(c) removal of  $\text{CO}$

(d) removal of  $\text{CH}_3\text{OH}$

on the equilibrium of the reaction :

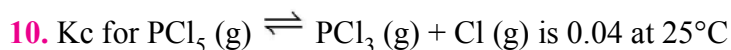
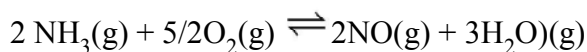




What will be the equilibrium constant for the decomposition of 1 mol of x ?



Determine the equilibrium constant for



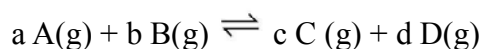
How much mole of  $PCl_5$  must be added to 3 L flask to obtain a chlorine concentration of 0.15 M ? [2.1]

### 3-Mark Questions

1. Four moles of  $PCl_5$  are heated in a closed 4 dm<sup>3</sup> container to reach equilibrium at 400 K. At equilibrium 50% of  $PCl_5$  is dissociated. What is the value of  $K_c$  for the dissociation of  $PCl_5$  into  $PCl_3$  and  $Cl_2$  at 400 K [0.50]

2. (a) How does the value of equilibrium constant predict the extent of a reaction ?  
(b) Equilibrium constant for a reaction is 10. What will be the equilibrium constant for the reverse reaction ?

3. (a) Define the equilibrium constant.  
(b) For the general reaction :

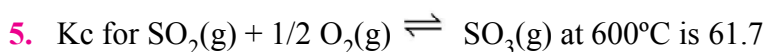


Derive the relationship between  $K_p$  and  $K_c$

4. (a) What is meant by dynamic nature of equilibrium ?  
(b) Consider the following transformation.

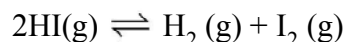


Calculate the value of  $K$  for  $A \rightleftharpoons D$



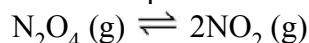
Calculate  $K_p$ . What is the unit  $K_p$  for the above equilibrium. [ $R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$ ]

6. 3.2 mole of HI were heated in a sealed bulb at  $444^\circ\text{C}$  till the equilibrium state was reached. Its degree of dissociation was found to be 20%. Calculate the number of moles hydrogen iodide, hydrogen and iodine present at the equilibrium point and also determine the equilibrium constant.



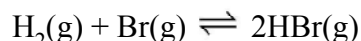
7. At  $25^\circ\text{C}$  and 1 atm, the partial pressure in an equilibrium mixture of  $\text{N}_2\text{O}_4$  and  $\text{NO}_2$  are 0.7 and 0.3 atm, respectively. Calculate the partial pressure when they are in equilibria at  $25^\circ\text{C}$  and at a total pressure of 10 atm.

8. 13.8 g of  $\text{N}_2\text{O}_4$  was placed in a 1L reaction vessel at 400 K and allowed to attain equilibrium



The total pressure at equilibrium was found to be 9.15 bar. Calculate  $K_c$ ,  $K_p$  and partial pressure at equilibrium. [ $K_c = 2.6$ ]

9. The equilibrium constant for the following reaction is  $1.6 \times 10^5$  at 1024 K



Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K. [NCERT]

[ $P_{\text{H}_2} \text{ eq.} = P_{\text{Br}_2} \text{ eq.} = 2.5 \times 10^{-2} \text{ bar}$ ;  $P_{\text{HBr}} = 10.0 \text{ bar}$ ]

10. In a reaction :

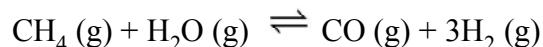


Indicate the direction in which equilibrium will shift when

- (i) Temperature is decreased.
- (ii) Pressure is decreased.
- (iii) Ne gas is added at constant pressure.

### 5-Mark Questions

1. Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction :



- (a) Write an expression for  $K_p$  for the above reaction.
- (b) How will the values of  $K_p$  and composition of equilibrium mixture be affected by :

- (i) increasing the pressure
- (ii) increasing the temperature
- (iii) using a catalyst ?

2. At 473 K, equilibrium constant  $K_c$  for decomposition of phosphorus pentachloride,  $\text{PCl}_5$  is  $8.3 \times 10^{-3}$ . If decomposition is depicted as,



$$\Delta H^\circ = 124.0 \text{ kJ mol}^{-1}$$

- (a) Write an expression for  $K_c$  for the reaction.
  - (b) What is the value of  $K_c$  for the reverse reaction at the same temperature?
  - (c) What would be the effect on  $K_c$  if (i) more  $\text{PCl}_5$  is added (ii) pressure is increased (iii) the temperature is increased ? [(b)  $K_c = 120.48$  (c) (i) No effect (ii) No effect (iii)  $K_c$  increase]
3. Ammonia is prepared by Haber's process in which the following reaction occurs



Mention the effect of following on the equilibrium concentration of ammonia.

- (a) Increasing pressure
  - (b) Increasing temperature
  - (c) Use of catalyst at an optimum temperature.
  - (d) Addition of inert gas at constant volume.
  - (e) Addition of inert gas at constant pressure.
4. Ethyl acetate is formed by the reaction between ethanol and acetic acid and the equilibrium is represented as :
- $$\text{CH}_3\text{COOH}(\text{l}) + \text{C}_2\text{H}_5\text{OH}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(\text{l}) + \text{H}_2\text{O}(\text{l})$$
- (i) Write the concentration ratio (reaction quotient),  $Q_c$ , for this reaction (note : water is not in excess and is not a solvent in this reaction)
  - (ii) At 293 K, if one starts with 1.00 mol of acetic acid and 0.18 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.
  - (iii) Starting with 0.5 mol of ethanol and 1.0 mol of acetic acid and maintaining it at 293 K, 0.214 mol of ethyl acetate is found after sometime. Has equilibrium been reached ?

[(ii) 3.92 (iii)  $Q_c = 0.204$ , Equilibrium has not been attained.]

## Ionic Equilibrium

### 1–Mark Questions

- Which of the following are Lewis acids ?  
 $\text{H}_2\text{O}$ ,  $\text{BF}_3$ ,  $\text{H}^+$ ,  $\text{NH}_4^+$
- Write the conjugate acids for the following Bronsted bases.  
 $\text{C}_6\text{H}_5\text{OH}$ ,  $\text{H}_2\text{O}$
- Write the conjugate bases for the following Bronsted acids.  
 $\text{H}_2\text{O}$ ,  $\text{CH}_3\text{COOH}$
- Which of the following are Lewis acids ?  
(a)  $\text{H}_2\text{O}$ , (b)  $\text{AlCl}_3$  (c)  $\text{NH}_4^+$
- Define the Ostwald's dilution law.
- $\text{SO}_3^{2-}$  is Bronsted base or acid and why ?
- Why pH of our blood remains almost constant at 7.4 though we quite often eat spicy food ?
- pH of black coffee is 5.0 at  $25^\circ\text{C}$ . Is black coffee acidic or basic ?
- What will be the value of  $(\text{pK}_a + \text{pK}_b)$  at  $25^\circ\text{C}$ .
- What will be the pH of 1 M  $\text{KNO}_3$  solutions at  $25^\circ\text{C}$ .
- $\text{CaCl}_2(\text{s}) + \text{sq} \rightleftharpoons \text{CaCl}_2(\text{aq}) + \text{Heat}$   
Discuss the solubility of  $\text{CaCl}_2$  if temperature is increased.
- Why does the solubility of  $\text{CO}_2$  decrease with rise in temperature ?
- The solubility of  $\text{A}_2\text{X}_3$  is  $y \text{ mol dm}^{-3}$ . Calculate its solubility product.
- Write the  $K_{\text{sp}}$  expression for  $\text{Al}(\text{OH})_3$
- What is the condition for precipitation of a salt ?
- Predict the nature of solution when  $\text{NH}_4\text{NO}_3$  undergo hydrolysis.
- Explain why pure  $\text{NaCl}$  precipitates out when  $\text{HCl}$  gas is passed through the solution of  $\text{NaCl}$  ?
- Give the Henderson's -Hasselbalch equation for an acidic buffer solution.
- On which of the factors the equilibrium depend :  
Temperature, nature of reactant and product, initial concentration and pressure of the reactants.

20. What are amphoteric substances ? Give one example.
21. What could be the temperature 15°C or 100°C for  $K_w = 7.5 \times 10^{-14}$ .
22. What happens to ionic product of water if some acid is added to it ?
23. What is the conjugate base of  $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$

### 2–Mark Questions

1. Explain ionic product of water. What is the effect of temperature on ionic product of water ?
2. What is pOH ? What is its value for neutral water at 25°C
3. (a) Define buffer solution  
(b) Give one example each of acidic and basic buffer
4. The solubility of  $\text{Ca}_3(\text{PO}_4)_2$  in water is x moles / litre. Calculate its solubility product. [108x<sup>5</sup>]
5. Calculate the pH of a 0.01 M solution of acetic acid.  $K_a$  for  $\text{CH}_3\text{COOH}$  is  $1.8 \times 10^{-5}$  at 25°C. [3.37]
6. Calculate pH when 9.8 g  $\text{H}_2\text{SO}_4$  is dissolved in 2 litre of solution [1]
7. Calculate the pH of  $10^{-9}$  M HCl. [pH = 6.9957]
8. An acid having pH = 6 is diluted 100 times. What will be the pH of the final solution ? [pH = 6.98]
9. Calculate the pH of  $10^{-10}$  M NaOH solution. [pH = 7.0004]
10. Solid  $\text{Ba}(\text{NO}_3)_2$  is gradually dissolved in a  $1.0 \times 10^{-4}$  M  $\text{Na}_2\text{CO}_3$  solution. At what concentration of  $\text{Ba}^{2+}$  will a precipitate begin to form ? ( $K_{sp}$  for  $\text{BaCO}_3 = 5.1 \times 10^{-9}$ ) [ $5.1 \times 10^{-5}$  M]
11. The p $K_a$  of acetic acid and p $K_b$  of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution. [7.005]

### 3–Mark Questions

1. What is meant by the conjugate acid-base pair ? Find the conjugate acid/base for the following species :  
 $\text{HNO}_2$ ,  $\text{CN}^-$ ,  $\text{HClO}_4$ ,  $\text{F}^-$ ,  $\text{OH}^-$ ,  $\text{OH}_3^-$ , and  $\text{S}^{2-}$
2. Define solubility product. Calculate the solubility product of  $\text{Ca}(\text{OH})_2$  if its solubility is 3. [K<sub>sp</sub> = 108]

3. Calculate the pH of the following mixture.  
200 ml of M/10  $\text{H}_2\text{SO}_4$  + 400 ml of M/10  $\text{H}_3\text{PO}_4$  + 400 ml of M/10 HCl  
[0.69]
4. The solubility product of AgCl in water is  $1.5 \times 10^{-10}$ . Calculate its solubility in 0.01 M NaCl aqueous solution. [1.5  $\times 10^{-8}$  M]
5. Predict whether a precipitate will be formed or not on mixing 20 ml of 0.001 M NaCl solution with 80 ml of 0.01 M  $\text{AgNO}_3$  solution,  $K_{sp}$  for AgCl is  $1.5 \times 10^{-10}$ . [Yes, ppt. will be formed]
6. Equal volumes of three acids solution with pH 3, 4 and 5 are mixed in a vessel. What will be the  $\text{H}^+$  ion concentration in the mixture ?
7. Calculate the pH of a solution obtained by mixing 50 ml of 0.2 M HCl and 50 ml of 0.1 M NaOH. [1.3010]
8. The pH of 0.1 M hydrocyanic acid solution is 5.2. What is the value of  $K_a$  for hydrocyanic acid ? [ $K_a = 3.69 \times 10^{-10}$ ]
9. A buffer solution is prepared by mixing equal concentration of weak base and its salt with strong acid. For the base  $K_b$  is given as  $10^{-9}$ . Calculate the pH of the buffer solution. [5]
10. Calculate the pH of the resultant mixtures :
  - (a) 10 mL of 0.2 M  $\text{Ca(OH)}_2$  + 25 mL of 0.1 M HCl
  - (b) 10 mL of 0.01 M  $\text{H}_2\text{SO}_4$  + 10 mL of 0.01 M  $\text{Ca(OH)}_2$
  - (c) 10 mL of 0.1 M  $\text{H}_2\text{SO}_4$  + 10 mL of 0.1 M KOH [NCERT]  
[(a) 12.632 (b) 7 (c) 1.3]
11. Equal volumes of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate ? (For cupric iodate  $K_{sp} = 7.4 \times 10^{-8}$ ). [NCERT] [NO ppt. will occur]
12. What is the maximum concentration of equimolar solutions of ferrous sulphate and sodium sulphide so that when mixed in equal volumes, there is no precipitation of iron sulphide ? (For iron sulphide,  $K_{sp} = 6.3 \times 10^{-18}$ ) [NCERT] [ $5.02 \times 10^{-9} \text{ mol L}^{-1}$ ]
13. The ionization constant of aniline is  $3.69 \times 10^{-10}$  :
  - (i) Calculate pH of 0.01 M solution of aniline.
  - (ii) Calculate the degree of dissociation of aniline in the solution.
  - (iii) Calculate the ionization constant of conjugate acid of the aniline.

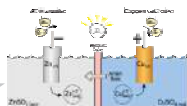
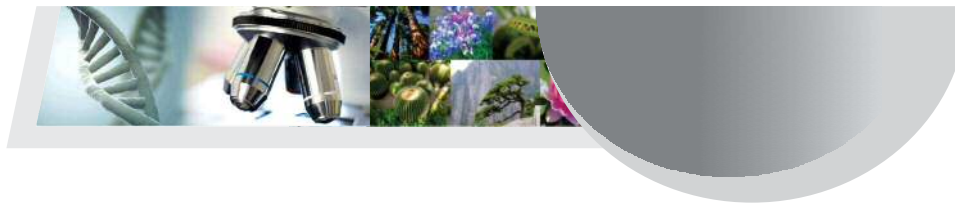


14. The  $K_{sp}$  for  $\text{Al}(\text{OH})_3$  is  $2.7 \times 10^{-11}$ . Calculate its solubility in g/L and also pH of this solution.
15. Calculate the volume of water required to dissolve 0.3 gm of  $\text{Zn}(\text{OH})_2$  to get a saturated solution.  $K_{sp} [\text{Zn}(\text{OH})_2] = 1 \times 10^{-15}$   
[Atomic mass of Zn = 65.3, O = 16, H = 1]
16. The pH of an aqueous solution of ammonia is 11.5. Find the molarity of the solution.  
 $K_b (\text{NH}_4\text{OH}) = 1.8 \times 10^{-5}$  [0.57]
17. What is the pH of the solution when 0.2 mole of HCl is added to 1 L of the solution containing 0.1 M each of  $\text{CH}_3\text{COOH}$  & acetate ion. Assume that volume is 1 L.  
 $K_a (\text{CH}_3\text{COOH}) = 1.8 \times 10^{-5}$  [1]

### 5-Mark Questions

1. Calculate the pH value of the following solutions.  
(a)  $10^{-2}$  M HCl.  
(b)  $10^{-3}$  M  $\text{H}_2\text{SO}_4$ .  
(c)  $10^{-4}$  M NaOH  
(d) 0.04 M NaOH.  
(e) 0.03 M HCl. [(a) 2 (b) 2.6990 (c) 10 (d) 12.60 (e) 1.5229]
2. (i) Calculate the pH of a buffer solution containing 2.0 mol/L  $\text{CH}_3\text{COOH}$  in 1.0 mol/L of  $\text{CH}_3\text{COONa}$ . [ $K_a$  for  $\text{CH}_3\text{COOH}$  is  $1.8 \times 10^{-5}$  at  $25^\circ\text{C}$ .]  
(ii) Calculate the change in pH of the buffer after the addition of 0.01 mol NaOH.  
(iii) Calculate the pH of the buffer after the addition of 0.01 M HCl ?  
[H. O. T.S] [(i) 4.4337, (ii) 0.0065, (iii) 4.4372]
3. (a) What is solubility product ? How is it different from ionic product ?  
(b) Calculate the solubility of  $\text{PbCl}_2$  if its solubility product is  $1.0 \times 10^{-6}$  at 298 K. [ $6.3 \times 10^{-3}$  M]
4. (a) Write short notes on :  
(i) Common ion effect  
(ii) Buffer capacity

- (b) Calculate the pH of a buffer solution containing 0.2 mole of  $\text{NH}_4\text{Cl}$  and 0.1 mole of  $\text{NH}_4\text{OH}$  per litre.  $K_b$  for  $\text{NH}_4\text{OH} = 1.8 \times 10^{-5}$  [8.966]
5. 500 ml saturated solution of  $\text{Ca}(\text{OH})_2$  is mixed with equal volume of 0.4 M  $\text{NaOH}$ . How much  $\text{Ca}(\text{OH})_2$  in mg is precipitated ?  
[ $K_a (\text{Ca}(\text{OH})_2) = 4.42 \times 10^{-5}$  at  $25^\circ\text{C}$ ] [H.O.T.S]
6. Calculate the pH of 0.1 M ammonia solution. Calculate the pH after 50 ml of this solution is treated with 25 ml of 0.1 M  $\text{HCl}$ .  
 $K_b (\text{NH}_3) = 1.77 \times 10^{-5}$  [11.12, 9.24]



## Chapter - 8

# Redox Reactions

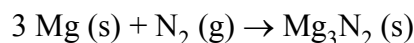
Oxidation	Reduction
1. Addition of oxygen	1. Removal of oxygen
2. Removal of an Hydrogen	2. Addition of Hydrogen
3. Addition of an electronegative element.	3. Removal of an electronegative element.
4. Addition of an electropositive element	4. Addition of an electropositive element.
5. Loss of electron(s)	5. Gain of electron(s)
6. Increase in oxidation number.	6. Decrease in oxidation, number.

- **Reducing Agent** : Donor of electron(s).
- **Oxidising Agent** : Acceptor of electron(s).
- **Redox Reaction** : Reactions in which oxidation and reduction takes place simultaneously.
- **Oxidation Number** : It is charge that an atom appears to have in a given species when the bonding electron are counted towards more electro-negative atom.
- **Calculation of Oxidation Number** :
  - O.S. of all the elements in their elemental form (in standard state) is taken as zero. O.S. of element of  $\text{Cl}_2$ ,  $\text{F}_2$ ,  $\text{O}_2$ ,  $\text{P}_4$ ,  $\text{O}_3$ ,  $\text{Fe(s)}$ ,  $\text{H}_2$ ,  $\text{N}_2$ ,  $\text{C}$  (graphite) is zero.
  - Common O.S. of elements of 1<sup>st</sup> group is +1. Common O. S. of elements of 2<sup>nd</sup> group + 2.
  - For ions composed of only one atom, the oxidation number is equal to the charge on the ion.

- (d) The oxidation number of oxygen in most compounds is  $-2$ . While in peroxides (*e.g.*,  $\text{H}_2\text{O}_2$ ,  $\text{Na}_2\text{O}_2$ ), each oxygen atom is assigned an oxidation number of  $-1$ , in super oxides (*e.g.*,  $\text{KO}_2$ ,  $\text{RbO}_2$ ) each oxygen atom is assigned an oxidation number of  $-\frac{1}{2}$ .
- (e) In oxygen di fluoride ( $\text{OF}_2$ ) and dioxygen difluoride ( $\text{O}_2\text{F}_2$ ), the oxygen is assigned an oxidation number of  $+2$  and  $+1$ , respectively.
- (f) The oxidation number of hydrogen is  $+1$  but in metal hydride its oxidation no. is  $-1$ .
- (g) In all its compounds, fluorine has an oxidation number of  $-1$ .
- (h) The algebraic sum of the oxidation number of all the atoms in a compound must be zero.
- (i) In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must equal the charge on the ion.

● **Types of Redox Reactions:**

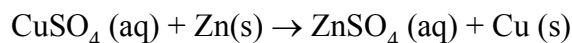
(i) **Combination Reaction :**       $0 \quad 0 \quad +2 \quad -3$



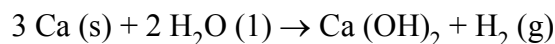
(ii) **Decomposition :**  $+1 +5 -2 \quad +1 -1 \quad 0$



(iii) **Metal Displacement :**  $+2 +6 -2 \quad 0 \quad +2 +6 -2 \quad 0$

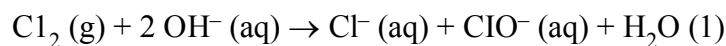


(iv) **Non-metal displacement :**  $0 \quad +1 -2 \quad +2 -2 +1 \quad 0$



(v) **Disproportionation :** It is a reaction in which same element is reduced and oxidized simultaneously.

$0 \quad -1 \quad +1$

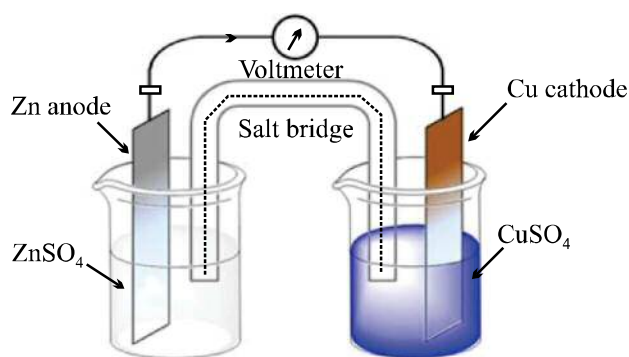


- **Stock Notation :** Representing oxidation number of metal in Roman numerals within parenthesis after the symbol or name of metal in the molecular formula or name of a compound. *e.g.*, Stock Notation of Ferric oxide is  $\text{Fe}_2\text{(III)O}_3$  or Iron (III) oxide.

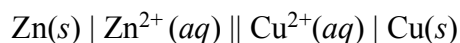
- **Fractional Oxidation Number :** When two or more atoms of an element are present in different oxidation states, then calculated oxidation number may comes out as fractional due to average of all the different oxidation states.

In reality no element can have a fractional oxidation state.. .

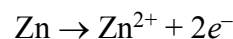
- **Electrode or Redox Couple** : A redox couple is defined as having together oxidized and reduced forms of a substance taking part in an oxidation or reduction half reaction.
- **Electrode Potential (E)** : Tendency of an electrode to gain or lose electrons.
- **Standard Electrode Potential (E°)** : Electrode Potential measured at 298 K and 1M concentration of metal ions (or 1 bar pressure of gas).
- **Electrochemical Cell** : A device in which chemical energy of a spontaneous redox reaction is converted into electrical energy.



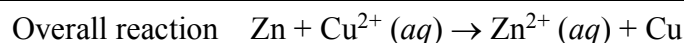
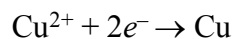
**Cell diagram,**



LHS oxidation,



RHS reduction



- Representation of an Electrochemical cell :

————— Flow of electrons —————→

←————— Flow of current —————

Left Electrode

Salt Bridge

Right Electrode

**Oxidation**

**Reduction**

Anode

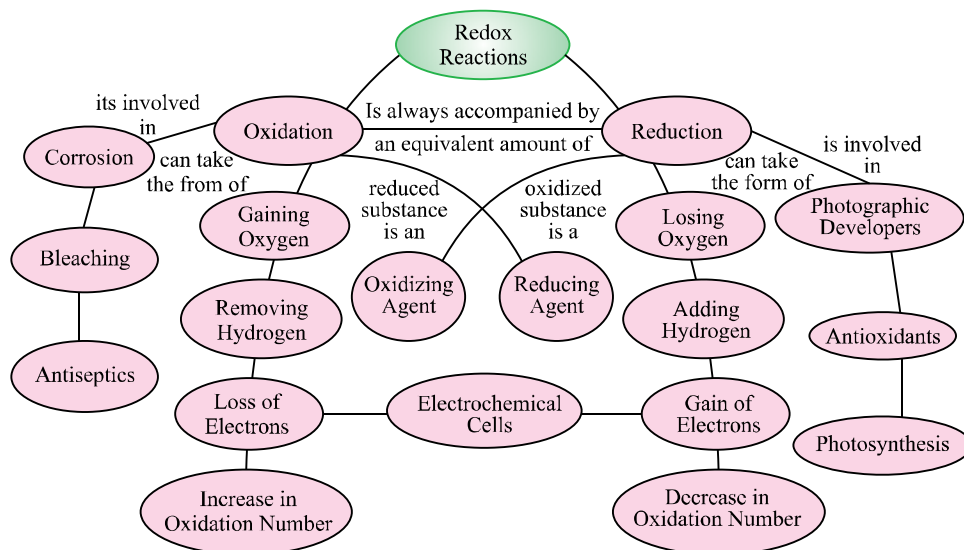
Cathode

Negative

Positive

- **Functions of Salt Bridge** : (i) To complete inner circuit. (ii) To maintain electrical neutrality.

## Flow learning



### 1-Mark Questions

#### Oxidation-Reduction : Classical, Electronic and Oxidation Number Concept

1. Define oxidation and reduction according to electronic concept.
2. Define oxidation and reduction according to oxidation number.
3. A freshly cut apple is almost white but it turns reddish brown after some. Give reason.
4. Define oxidation number.
5. Write oxidation number of Mn in  $\text{KMnO}_4$ .
6. Write oxidation Number of Cr in  $\text{Cr}_2\text{O}_7^{2-}$ .
7. Write Stock Notation of  $\text{MnO}_2$  and  $\text{AuCl}_3$ .

#### Redox Reactions

8. Define redox reaction with example.
9. Define disproportionation reaction. Give one example.
10. Define the term Redox titration.
11. Name the indicator used in redox titrations involving  $\text{K}_2\text{Cr}_2\text{O}_7$  as an oxidizing agent.

#### Redox Reactions and Electrode Processes

12. At what concentration of  $\text{Cu}^{2+}(\text{aq})$  will electrode potential become equal to its standard electrode potential ?  
[Ans. 1 M]

13. The standard reduction potentials of three metals cations X, Y and Z are + 0.52, - 3.03 and - 1.18 V respectively. Arrange X, Y and Z in order of increasing reducing power. [Ans.  $X < Z < Y$ ]
14. An electrochemical cell consists of two electrodes *i.e.*, Anode and Cathode. What is the direction of flow of electrons in this cell ?
15. Why anode is negatively charged in an electrochemical cell.
16. Out of Zn and Cu vessel one will be more suitable to store 1 M HCl ?

[Ans. Cu]

Given  $E_{\text{Zn}^{2+}/\text{Zn}}^{\circ} = -0.76 \text{ V}$ ,  $E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} = +0.34 \text{ V}$ .

15. Is it safe to stir 1 M  $\text{AgNO}_3$  solution with copper spoon ? [Ans. No]

Given  $E_{\text{Ag}^+/\text{Ag}}^{\circ} = +0.80 \text{ V}$ ,  $E_{\text{Cu}^{2+}/\text{Cu}}^{\circ} = +0.34 \text{ V}$ .

## 2 Mark Questions

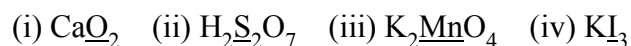
### Oxidation-Reduction: Classical, Electronic and Oxidation Number Concept

- Identify oxidant and reductant in the reaction :  $\text{I}_2 (\text{aq}) + 2\text{S}_2\text{O}_3^{2-} (\text{aq}) \rightarrow 2\text{I}^- (\text{aq}) + \text{S}_4\text{O}_6^{2-} (\text{aq})$ .
- Calculate oxidation number of Fe in  $\text{Fe}_3\text{O}_4$  and write a suitable justification of your answer.
- Oxidation-reduction reactions are complementary. Explain.
- Write formula for the following compounds :
  - Mercury (II) chloride
  - Nickel (II) sulphate
  - Iron (III) sulphate
  - Chromium (III) oxide

## Redox Reactions

- Justify that the reaction :  $\text{H}_2\text{O}(\text{s}) + \text{F}_2 \rightarrow \text{HF} + \text{HOF}$  is a redox reaction. [NCERT]
- A decomposition reaction may or may not be a redox reaction. Write two decomposition reactions in support of the statement.
- Split the reaction  $2\text{K} (\text{s}) + \text{Cl}_2 (\text{g}) \rightarrow 2\text{KCl} (\text{s})$  into oxidation and reduction half reactions.

8. Calculate the oxidation number of underlined elements in following compounds :



### Redox Reactions and Electrode Processes

9. Write the functions of salt bridge in an electrochemical cell.
10. Define the term redox couple. Write the practical application of redox couple.
11. The standard reduction potentials of two metals A and B are  $-0.76\text{ V}$  and  $+0.34\text{ V}$  respectively. An electrochemical cell is formed using electrodes of these metals.
- (i) Identify the cathode and anode.
- (ii) Write the direction of flow of electron.

### 3 Mark Questions

#### Oxidation-Reduction : Classical, Electronic and Oxidation Number Concept

1. Calculate oxidation number of :
- (i) Cr in  $\text{Cr}_2\text{O}_4^{2-}$
- (ii) O in  $\text{KO}_2$
- (iii) Na in  $\text{Na}_2\text{O}_2$ .
2. Account for the following :
- (i)  $\text{HNO}_3$  acts as oxidizing agent while  $\text{HNO}_2$  can act both as reducing and oxidizing agent.
- (ii)  $\text{AgF}_2$  is unstable compound and a strong oxidizing agent.
- (iii) Ozone acts as an oxidizing agent.

### Redox Reactions

3. Permanganate ion ( $\text{MnO}_4^-$ ) reacts with sulphur dioxide gas in acidic medium to produce  $\text{Mn}^{2+}$  ion and hydrogen sulphate ion. Write ionic equation and balance by ion electron method.
4. Balance the following equation by oxidation number method :
- $$\text{P}_4(\text{s}) + \text{OH}^-(\text{aq}) \rightarrow \text{PH}_3 + \text{H}_2\text{PO}_2^-(\text{aq})$$
5. Balance the following equation in basic medium :
- $$\text{Cl}_2\text{O}_7(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{ClO}_2^-(\text{aq}) + \text{O}_2(\text{g})$$



### Redox Reactions and Electrode

6. Depict the galvanic cell in which the reaction  $\text{Zn (s)} + 2 \text{Ag}^+ \text{(aq)} \rightarrow \text{Zn}^{2+} \text{(aq)} + 2 \text{Ag (s)}$  takes place. Further show :
- (i) Which electrode is negatively charged ?
  - (ii) The carriers of the current in the cell
  - (iii) Individual reaction at each electrode. [NCERT]
7. Explain why ?
- (i) Reaction  $\text{FeSO}_4 \text{(aq)} + \text{Cu (s)} \rightarrow \text{CuSO}_4 \text{(aq)} + \text{Fe}$  does not occur.
  - (ii) Zinc can displace copper from aqueous  $\text{CuSO}_4$  solution but Ag cannot.
  - (iii) Solution of  $\text{AgNO}_3$  turns blue when copper rod is immersed in it.

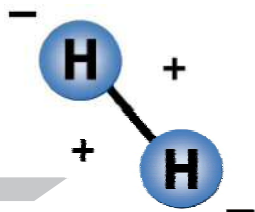
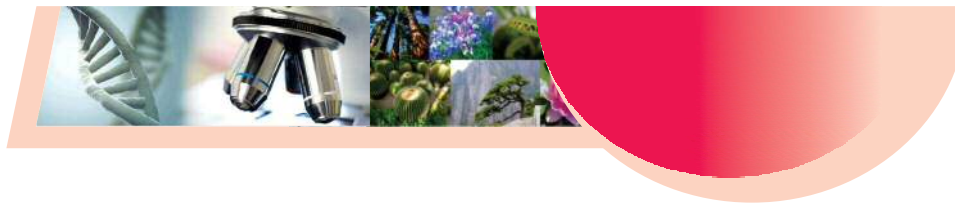
### 5 Mark Questions

### Redox Reactions

1. (i)  $\text{MnO}_4^{2-}$  undergoes disproportionation reaction in acidic medium but  $\text{MnO}_4^-$  does not. Give reason.
- (ii) Give one example each of the following redox reactions:
- (a) Combination reaction
  - (b) Decomposition reaction
  - (c) Metal displacement reaction

### Redox Reactions and Electrode Processes

2. Consider the cell reaction of an electrochemical cell :  $\text{Ni(s)} + 2 \text{Ag}^+ \text{(aq)} \rightarrow \text{Ni}^{2+} \text{(aq)} + 2 \text{Ag (s)}$  and answer the following questions :
- (i) Write anode and cathode half reactions.
  - (ii) Mention the direction of flow of electrons.
  - (iii) How is the electrical neutrality maintained in the solutions of the two half cells.
  - (iv) Write the formula for calculating standard emf of this cell.
  - (v) How does the emf change when the concentration of silver ions is decreased ?



## Chapter - 9

# Hydrogen

### Part A

- Hydrogen is the first element in the periodic table and also the lightest element known.
- Isotopes of hydrogen :**
  - Protium ( ${}_1^1\text{H}$ )
  - Deuterium ( ${}_1^2\text{H}$ )
  - Tritium ( ${}_1^3\text{H}$ )
- Preparation of Dihydrogen :**
  - Laboratory preparation :  $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$ .
  - Commercial preparation : By electrolysis of acidified water.
  - High purity dihydrogen is obtained by electrolysis of warm aqueous barium hydroxide.
- Water gas :** The mixture of  $\text{CO} + \text{H}_2$  is known as **water gas**. It is also known as '**Syn**' Gas.
- Water gas shift reaction :**

$$\text{CO(g)} + \text{H}_2\text{O(g)} \xrightarrow[\text{Catalyst}]{673 \text{ K}} \text{CO}_2\text{(g)} + \text{H}_2\text{(g)}$$
- Properties of Dihydrogen :**

The H-H bond dissociation enthalpy is highest for a single bond dissociation enthalpy between two atoms of any element.

It is relatively inert at room temperature due to the high H-H bond enthalpy.
- Uses of Dihydrogen :**
  - For synthesis of Ammonia ( $\text{NH}_3$ )

(ii) For production of Methanol ( $\text{CH}_3\text{OH}$ )

(iii) In oxyhydrogen torches

(iv) In a fuel cell

- **Hydrides**

(i) Ionic or salt like or saline hydrides are formed with most of the *s*-block elements. Significant covalent character is found in  $\text{LiH}$ ,  $\text{BeH}_2$  and  $\text{MgH}_2$ .

- Covalent or Molecular hydrides are formed with most of the *p*-block elements. There are further classified as :

(a) **Electron deficient hydrides** are formed by group 13 elements *e.g.*,  $\text{B}_2\text{H}_6$ . These acts as Lewis acid.

(b) **Electron Precise hydrides** are formed by group 14 elements *e.g.*,  $\text{CH}_4$ .

(c) **Electron rich hydrides** have lone pair of electrons on central atoms of the molecules. Elements of group 15-17 form these types of hydrides.  $\text{NH}_3$ ,  $\text{HF}$  has high m.p./b.p. due to presence of intermolecular hydrogen bonding.

(iii) **Metallic or Non-stoichiometric or Interstitial hydrides** are formed by *d* and *f* block elements. For example  $\text{La H}_{2.87}$ .

These are potential source of hydrogen.

## Part-B

- **Water : ( $\text{H}_2\text{O}$ )**

**Hard water :** Hard water contains calcium and magnesium salts in the form of hydrogencarbonate, chloride and sulphate. Hard water does not give lathers with soap.

**Soft water :** Water free from soluble salts of calcium and magnesium is soft water.

**Types of Hardness :**

**Temporary hardness** is due to presence of calcium or magnesium hydrogen carbonate in water.

Temporary hardness can be removed by :

(i) Boiling

(ii) Clark's Method

**Permanent hardness :**

Such hardness is due to presence of calcium or magnesium chlorides and sulphates.

Permanent hardness can be removed by :

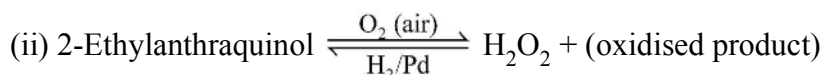
- (i) Calgon's method
- (ii) Ion exchange method.

**Demineralised or Deionised water :** Water free from all soluble mineral salts is known as **demineralised water**.

- **Hydrogen Peroxide (H<sub>2</sub>O<sub>2</sub>)**

**Preperation :**

- (i) By electrolytic oxidation of acidified sulphate solutions at high current density.



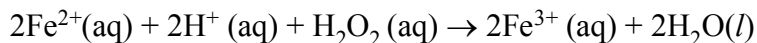
- **Physical Properties**

- (i) Miscible with water in all proportions.
- (ii) A 30% of H<sub>2</sub>O<sub>2</sub> solution is marketed as '100 volume of hydrogen peroxide'.

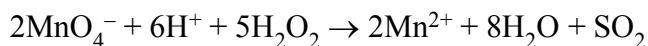
- **Chemical Properties :**

- (i) It acts as an oxidising as well as reducing agent.

- (ii) **Oxidising action in acidic medium :**



- (iii) **Reducing action in acidic medium :**



- **Storage of H<sub>2</sub>O<sub>2</sub> :**

- (i) Stored in wax-linked glass or plastic vessels in dark. Urea can be added as a stabiliser.

- (ii) It is kept away from dust.

- **Uses of H<sub>2</sub>O<sub>2</sub> :**

- (i) As an antiseptic it is sold in the market name perhydrol.
- (ii) In synthesis of hydroquinone.
- (iii) As a bleaching agent.

## Part-A

### 1-Mark Questions

1. Name the isotope of hydrogen which is radioactive in nature. [Ans. Tritium]

2.  $\text{H}^+$  ions does not exist freely and is always associated with other atoms or molecule. Explain.
3. Give the composition of water gas. [Ans. CO,  $\text{H}_2$ ]
4. Name the compound whose electrolysis in aqueous state, give high purity (99.95%) dihydrogen. [Ans. aq  $\text{Ba}(\text{OH})_2$  solution]
5. Give the main purpose of water gas shift reaction.
6. Write the chemical reaction occurring during coal gasification.
7. Name the element used in fuel cell for generating electricity. [Ans.  $\text{H}_2$ ]
8. Give an example of electron deficient covalent hydride. [Ans.  $\text{B}_2\text{H}_6$ ]
9. Name the hydrides which have high potential for hydrogen storage. [Ans. Metallic hydrides]
10. Name the groups in *d*-block elements which do not form metallic hydrides. [Ans. 7, 8, 9]
11.  $\text{H}_2$  is relatively inert at room temperature. Explain.
12. Complete the reaction :  

$$\text{C(s)} + \text{H}_2\text{O(g)} \xrightarrow{1270\text{ K}} \text{CO(g)} + \text{H}_2\text{(g)}$$
13. Name the phenomenon as a reason of which water has unusual boiling point. [Ans. Extensive hydrogen bonding]
14. Draw structure of water.
15. At atmospheric pressure ice crystallised in the ..... form but at very low temperature it condenses to ..... form. [Ans. Hexagonal, cubic]
16. Mention the temperature at which density of ice is maximum. [Ans.  $4^\circ\text{C}$ ]
17. Density of ice ..... than density of liquid water. [Ans. Less]
18. Complete the reaction :  

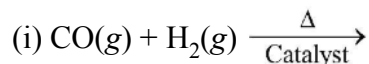
$$2\text{H}_2\text{O(l)} + 2\text{Na(s)} \longrightarrow$$
19. How many hydrogen-bonded water molecules (*s*) are associated in  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ . [Ans. One]
20. Name the compound used in Clark's method to remove temporary hardness of water. [Ans. Lime]
21. Write the chemical formula of "Calgon". [Ans.  $\text{Na}_4\text{P}_6\text{O}_{18}$ ]
22. A 30% solution of  $\text{H}_2\text{O}_2$  is marketed as ..... volume. [Ans. 100 volume]

23. Draw gas phase structure of  $\text{H}_2\text{O}_2$ .
24. Name the organic compound whose auto-oxidation is used to produce  $\text{H}_2\text{O}_2$  commercially or industrially. [Ans. 2-Ethylanthraquinol]

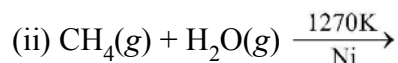
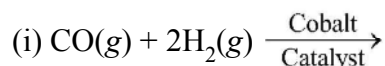
### Part-A

#### 2-Mark Questions

1. Complete the following reactions :



2. Among  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  and  $\text{HF}$  which would you expect to have highest magnitude of hydrogen bonding and why ?
3. How do you expect the metallic hydrides to be useful for hydrogen storage? Explain.
4. How can the production of dihydrogen obtained from “Coal gasification” can be increased ?
5. Write the name of isotopes of hydrogen. What is the mass ratio of these isotopes ?
6. Complete the reactions :



7. Comment on the reactions of dihydrogen with :

(i) Chlorine, (ii) Sodium.

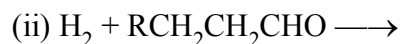
8. Arrange the following :

(i)  $\text{LiH}$ ,  $\text{NaH}$ ,  $\text{CsH}$  (In increasing order of ionic character)

(ii)  $\text{H}-\text{H}$ ,  $\text{D}-\text{D}$ ,  $\text{F}-\text{F}$  (In decreasing order of bond dissociation enthalpy)

9. Give two uses of dihydrogen.

10. Complete the reactions :



## Part-B

11. Give two reactions to show photeric nature of water.
12. Complete the reactions :
  - (i)  $2\text{F}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \longrightarrow$
  - (ii)  $6\text{CO}_2(\text{g}) + 12\text{H}_2\text{O}(\text{l}) \longrightarrow$
13. What is the difference between the term hydrolysis and hydration.
14. What do you understand by term 'autoprotolysis' of water ? What is its significance ?
15. What causes the temporary and permanent harness of water ?
16. Is demineralised or distill water useful for drinking purposes ? If not, how can it be made useful ?
17. Explain the terms : (i) Hydrogen economy. (ii) Fuel cell.
18. Write chemical reactions to justify that hydrogen peroxide can function as an oxidising as well as reducing agent.
19. Compare the structure of  $\text{H}_2\text{O}$  and  $\text{H}_2\text{O}_2$ .
20. How does  $\text{H}_2\text{O}_2$  behaves as a bleaching agent ?

## Part-A

### 3-Mark Questions

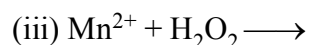
1. Complete the chemical reactions :
  - (i)  $8\text{LiH} + \text{Al}_2\text{Cl}_6 \longrightarrow$
  - (ii)  $2\text{LiH} + \text{B}_2\text{H}_6 \longrightarrow$
2. What do you understand by : (i) electron deficient, (ii) Electron precise, (iii) Electron rich compounds of hydrogen ? Provide justifications with suitable examples.
3. What do you understand by the term "non stoichiometric hydrides" ? Do you expect this type of the hydrides to be formed by alkali metals. Explain and Justify your answer.
4. Arrange the following :
  - (i)  $\text{CaH}_2$ ,  $\text{BeH}_2$ ,  $\text{TiH}_2$  (in order of increasing electrical conductance)
  - (ii)  $\text{NaH}$ ,  $\text{MgH}_2$ ,  $\text{H}_2\text{O}$  (in order of increasing bond dissociation enthalpy)
  - (iii)  $\text{Li}$ ,  $\text{F}$ ,  $\text{H}$  (in order of increasing ionisation enthalpy)

5. What do you understand by the terms :
- (i) Syn gas
  - (ii) Water gas shift reaction
  - (iii) Producer gas.
6. Would gas except the hydrides of N, O and F to have lower boiling point than the hydrides of their subsequent group members ? Give reasons.
7. Can phosphorous with outer electronic configuration  $3s^23p^3$  form  $\text{PH}_5$  ? Explain.
8. Why and how the hydrogen is regarded as a fuel of future ? Explain.
9. Write the reactions when dihydrogen reacts with (i)  $\text{O}_2$  (ii)  $\text{N}_2$  (iii)  $\text{Cl}_2$  under specific conditions.
10. Name the hydrides :
- (i) Which is non stoichiometric in nature ?
  - (ii) Which are stoichiometric compounds ?
  - (iii) Which has electron rich type hydrides ?

### Part-B

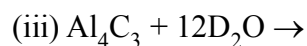
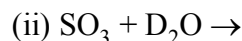
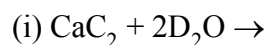
11. Complete the reactions :
- (i)  $\text{CaO}(s) + \text{H}_2\text{O}(g) \longrightarrow$
  - (ii)  $\text{AlCl}_3(g) + \text{H}_2\text{O}(l) \longrightarrow$
  - (iii)  $\text{Ca}_3\text{N}_2(s) + \text{H}_2\text{O}(l) \longrightarrow$
12. Discuss the principle and method of softening of hard water by synthetic exchange of resin method.
13. What is meant by 'demineralised' water and how can it be obtained ?
14. What properties of water make it useful as a solvent ? What types of compound can it (i) dissolved (ii) hydrolyse ?
15. Calculate the strength of 10 volume solution of  $\text{H}_2\text{O}_2$ .
16. Complete the reactions :
- (i)  $2\text{Fe}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \longrightarrow$
  - (ii)  $\text{HOCl} + \text{H}_2\text{O}_2 \longrightarrow$





17. Give three uses of  $\text{H}_2\text{O}_2$ .

18. Complete the reactions :



19. Give the limitations of using  $\text{H}_2$  as a fuel.

20.  $\text{H}_2\text{O}_2$  is stored in a wax lined glass or plastic vessels. Explain an equation showing decomposition of  $\text{H}_2\text{O}_2$  on exposure to light.

### Part-C

#### 5-Mark Questions

1. Answer the following :

(a) Name the most abundant form of hydrogen isotope. [Ans.  ${}_1^1\text{H}$ ]

(b) Name the particles emitted by tritium. [Ans.  $\beta^-$ ]

(c) Mixture of CO and  $\text{H}_2$  is used for preparation ..... [Ans. Methanol]

(d) Name the catalyst used in Haber's Process for manufacture of  $\text{NH}_3(\text{g})$ .  
[Ans. Fe]

(e) Name two electron rich hydrides. [Ans.  $\text{NH}_3$ ,  $\text{H}_2\text{O}$ ]

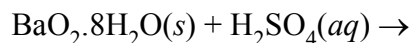
### Part-B

2. Answer the following :

(a) During Clark's method. Name the compound in which Mg is precipitated out. [Ans. Magnesium Hydroxide]

(b) Give the formula of Zeolite used in ion exchange method to remove permanent hardness of water. [Ans.  $\text{NaAlSiO}_4$ ]

(c) Complete the reaction :



(d)  $\text{H}_2\text{O}_2$  is miscible with water. Assign reason.

(e) Name the compound when can be used as a hair beach, mild antiseptic in the form of perhydrol. [Ans.  $\text{H}_2\text{O}_2$ ]



## Chapter - 10

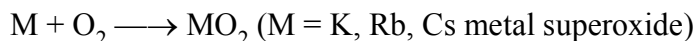
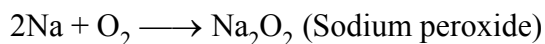
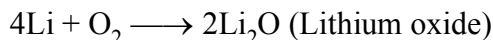
# The s-Block Elements

- s-Block elements consists of group-I (Alkali metals) and group-2 (Alkaline earth metals).
- Group 1<sup>st</sup> elements—Li, Na, K, Rb, Cs, Fr.
- Group 2<sup>nd</sup> elements—Be, Mg, Ca, Sr, Ba, Ra.
- **Atomic radius**—Atomic radius of alkali metals are greater than alkaline earth metals.
- **Hydration enthalpy** : Decreases with increases in ionic sizes.
- **Ionic mobility** : Smaller the size of ion, more highly it is hydrated and hence lower is its ionic mobility.



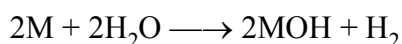
- **Ionisation enthalpies** : 1<sup>st</sup> I.E. of group 1<sup>st</sup> is smaller than group 2<sup>nd</sup> elements but 2<sup>nd</sup> I.E. of group 2<sup>nd</sup> is smaller than group 1<sup>st</sup> elements.
- **Flame colouration** : Due to low I.E., s-block elements and their salts imparts characteristics colour of oxidising flame (except Be and Mg). Be and Mg do not show flame colouration because they have small size and very high ionisation enthalpy.
- **Reducing character** : Due to large negative electrode potentials alkali metals are stronger reducing agent than alkaline earth metal.

- **Reactivity towards air** :

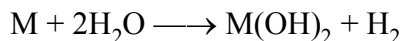


Alkaline earth metals being smaller in size do not form superoxides.

- **Reactivity towards H<sub>2</sub>O** :

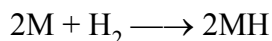


(Alkali metal)

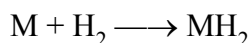


(Alkaline earth metals)

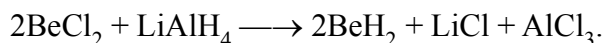
- **Reactivity towards hydrogen :**



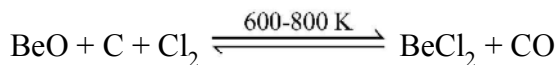
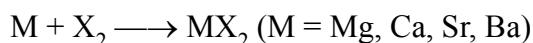
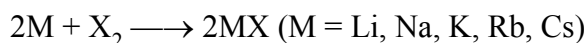
(M = Li, Na, K, Rb, Cs)



(M = Mg, Ca, Sr, Ba)

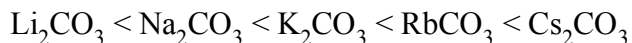


- **Reactivity towards halogens :**



- **Solution in liquid ammonia :** The fresh solution of alkali metals and alkaline earth metals (except Be and Mg) is deep blue, paramagnetic and highly reducing due to presence of ammoniated electrons.

- **Solubility of alkaline earth metal hydroxide in water :**



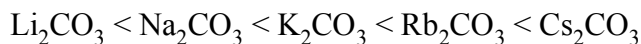
- **Solubility of alkaline earth metal carbonates in water.**



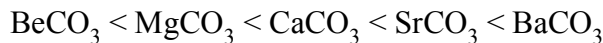
- **Solubility of alkaline earth metal sulphates in water :**



- **Thermal stability of alkali metal carbonates :**



- **Thermal stability of alkaline earth metal carbonates :**

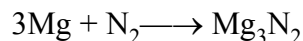
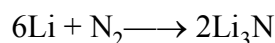


- **Anomalous behaviour of Li and Be :** It is due to very small size, high I.E. and high polarising power (*i.e.*, charge/radius)

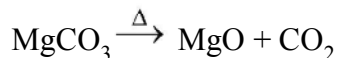
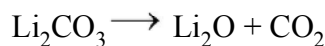
- **Diagonal relationship (similarities) between Li and Mg :**

(i) Both Li and Mg are hard.

(ii) Both react with  $N_2$  to form nitrides.

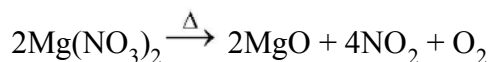
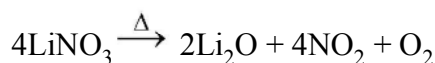


(iii) Decomposition of carbonates :



(iv) Both  $\text{LiCl}$  and  $\text{MgCl}_2$  are deliquescent. They form hydrates salts  $\text{LiCl} \cdot 2\text{H}_2\text{O}$  and  $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ .

(v) Decomposition of nitrates :



- Diagonal relationship (similarities) between Be and Al :

(i) Both are passive to acids due to formation of oxide layer.

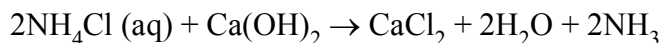
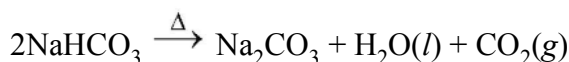
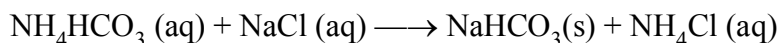
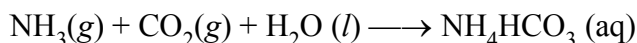
(ii) Hydroxides of both dissolve in alkali to form  $[\text{Be}(\text{OH})_4]^{2-}$  and  $[\text{Al}(\text{OH})_4]^-$ .

(iii) Chloride of both has bridged structure.

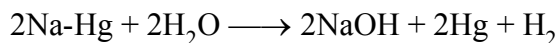
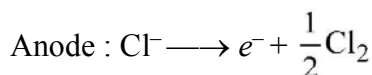
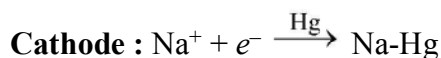
(iv) Both have tendency of form complexes of  $\text{BeF}_4^{2-}$ ,  $\text{AlF}_6^{3-}$ .

- Manufacturing of washing soda ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ) :

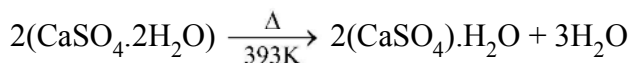
**Solvay process :**



- **Manufacturing of caustic soda ( $\text{NaOH}$ ) :** Castner-Kellner cell.



- **Plaster of paris :** ( $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$ )



Gypsum

- **Cement** is a finely powdered mixture of calcium silicates and aluminate along with small quantities of gypsum which sets into a hard stone like mass when treated with water.

## Group 1 Elements and their Compounds

### 1-Mark Questions

1. What is the oxidation state of K in  $\text{KO}_2$  ?
2. Why are group I element called alkali metals ?
3. Potassium carbonate cannot be prepared by solvay process. Why ?
4.  $\text{LiCl}$  is soluble in organic solvent. Why ?
5. Why are group 1 elements called alkali metals ?
6. Alkali metals are strong reducing agents. Why ?
7. Why do alkali metals give characteristics flame colouration ?
8. Arrange the following in order of increasing covalent character :  $\text{MCl}$ ,  $\text{MBr}$ ,  $\text{MF}$ ,  $\text{MI}$  (where  $\text{M}$  = Alkali metal) [**Ans.**  $\text{MF} < \text{MCl} < \text{MBr} < \text{MI}$ ]
9. Alkali metals can not be obtained by chemical reduction method. Explain.
10. Why is sodium metal kept under kerosene oil ?

## Group 2 Elements and their Compounds

11. Why Be and Mg do not give characteristics colour to the flame ?
12. Arrange the alkaline earth metal carbonate in the decreasing order of thermal stability.
13. Why do alkaline earth metals not form any superoxide ?
14. Why gypsum is added to cement ?
15. How plaster of paris is obtained from gypsum ?
16.  $\text{BeO}$  is insoluble in water but  $\text{BeSO}_4$  is soluble in water ? Why ?
17. Why second I.E. of group 2 elements is less than group 1 elements ?
18. What is quick lime ? How is it prepared ?
19. Why does Be shows similarities with Al ?
20. Name the alkaline earth meta hydroxide which is amphoteric.

## Group 1 Elements and their Compounds

### 2-Mark Questions

1. Why are alkali metals soft and have low melting points ?
2. Write any four similarities between Li and Mg. [NCERT]
3. Why are potassium and caesium rather than Lithium used in photoelectric cells ? [NCERT]
4. Why is  $\text{Li}_2\text{CO}_3$  decomposed at a lower temperature whereas  $\text{Na}_2\text{CO}_3$  at higher temperature ? [NCERT]
5. Among the alkali metals which has :
  - (i) Highest melting point.
  - (ii) Most electropositive character
  - (iii) Lowest size of ion.
  - (iv) Strongest reducing character.[Ans. (i) Li (ii) Cs (iii) Li (iv) Li]

## Group 2 Elements and their Compounds

6. Why does the solubility of alkali earth metal carbonates and sulphates decrease down the group ? [NCERT]
7. Draw the structure of  $\text{BeCl}_2$  in (i) Vapour phase (ii) Solid state. [NCERT]
8. When  $\text{CO}_2$  gas is passed in lime water it turns milky but in case of excess  $\text{CO}_2$  milkiness disappears. Support the statement by giving suitable reaction equations.
9. (i)  $E^\circ$  for  $\text{M}^{2+}(\text{aq}) + 2e^- \longrightarrow \text{M}(\text{s})$  (where  $\text{M} = \text{Ca}, \text{Sr}, \text{Ba}$ ) is nearly constant. [NCERT]  
(ii) What is dead burnt plaster ? How is it obtained from gypsum.
10. Write two important uses of (i) Limestone (ii) Quick lime.

## Group 1 Elements and their Compounds

### 3-Mark Questions

1. Assign reason for the following :
  - (i) Compounds of lithium are generally covalent.
  - (ii) Alkali metals are strong reducing agent.

- (iii) LiCl is more covalent than NaCl.
2. Discuss the various reactions that occur in Solvay process. [NCERT]
3. Explain why ?
- (i) Lithium salts are commonly hydrated.
- (ii) Sodium peroxide is widely used as an oxidising agent.
- (iii) Sodium wire is used to remove moisture from benzene but can't be used for drying alcohol.
4. Sodium hydroxide is generally prepared by electrolysis of brine solution in the Castner-Keller cell :
- (i) Write the reactions that occur in the cell.
- (ii) Write any two uses of NaOH.
5. State as to why ?
- (a) A solution of  $\text{Na}_2\text{CO}_3$  is alkaline.
- (b) Alkali metals are prepared by electrolysis of their fused chlorides.
- (c) Sodium is found to be more useful than potassium ?

### Alkaline Earth Metals and their Compounds

6. Arrange the following in order of property mentioned against each :
- (i)  $\text{BaCl}_2$ ,  $\text{MgCl}_2$ ,  $\text{BeCl}_2$ ,  $\text{CaCl}_2$  (Increasing ionic character)
- (ii)  $\text{Mg}(\text{OH})_2$ ,  $\text{Sr}(\text{OH})_2$ ,  $\text{Ba}(\text{OH})_2$ ,  $\text{Ca}(\text{OH})_2$  (Increasing solubility in water)
- (iii)  $\text{BeO}$ ,  $\text{MgO}$ ,  $\text{BaO}$ ,  $\text{CaO}$  (Increasing basic strength)
7. What happens when :
- (i) Mg is burnt in air.
- (ii) Quick lime is heated with silica.
- (iii) Chlorine is heated with slaked lime.
8. Write the raw material required for the manufacture of portland cement ? Why gypsum is added into it.
9. (i) Why alkaline earth metals cannot be obtained by reduction of their oxide ?
- (ii) Why the elements of group 2 are known as alkaline earth metals ?
10. (i) Alkaline earth metals form ionic salt having bivalent cations. Explain. Why ?
- (ii) A piece of magnesium ribbon continues to burn in  $\text{SO}_2$ . Why ?

## Group 1 Elements and their Compounds

### 5-Mark Questions

1. Explain the following observation :
- (a) LiI is more soluble than KI in ethanol.
  - (b) Sodium reacts with water less vigorously than potassium.
  - (c) LiF is insoluble in water.
  - (d) The mobilities of the alkali metal ions in aqueous solution are  $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$ .
  - (e) Lithium is the only alkali metal to form a nitride directly.

## Group 2 Elements and their Compounds

2. Complete the following reaction equations :
- (i)  $\text{BeCl}_2 + \text{LiAlH}_4 \rightarrow$
  - (ii)  $\text{CaO} + \text{SiO}_2 \rightarrow$
  - (iii)  $\text{Ca(OH)}_2 + \text{Cl}_2 \rightarrow$
  - (iv)  $\text{CaO} + \text{P}_4\text{O}_{10} \rightarrow$
  - (v)  $\text{C(OH)}_2 + \text{CO}_2 \rightarrow$





## Chapter - 11

# The p-Block Elements

### Group 13 Elements

- Electronic configuration :  $ns^2np^1$ .
- **Oxidation states** : B and Al show an oxidation state of + 3 only while Gallium, Indium and Thallium show oxidation states of both + 1 and + 3. Further due to inert pair effect, as we move down the group, the stability of + 3 oxidation state decreases while that of + 1 oxidation state increases.
- **Inert pair effect** : It is the reluctance of the  $s$ -electrons of the valence shell to take part in bonding. It arises due to poor or ineffective shielding of the  $ns^2$  electrons of the valence shell by the intervening  $d$  or  $f$  electrons. Inert pair effect increases down a group and thus the elements present in the lower part of the group show lower oxidation states which are two units less than the highest group oxidation state.
- **Halides** : All the elements of group 13 (except Tl) form trihalides of general formula  $MX_3$  ( $X = F, Cl, Br, I$ ).  
Boron trihalides exist as only monomers, whereas aluminium trihalides exist as dimers.
- **Nature of Trihalides** : The trihalides of group 13 elements have only six electrons in the valence shell and have a strong tendency to accept a pair of electrons to acquire the nearest inert gas configuration. So trihalides of group 13 elements behave as Lewis acids.
- **Borax** : It is  $Na_2B_4O_7 \cdot 10H_2O$ .
- **Orthoboric acid ( $H_3BO_3$ )** : It is a weak monobasic acid.
- **Diborane ( $B_2H_6$ )** : It is a dimer of  $BH_3$ .

## Group 14 Elements

- Electronic configuration :  $ns^2 np^2$
- **Oxidation states** : Group 14 elements shows + 2 and + 4 oxide state. As we move down the group stability of + 4 oxidation state decreases while of + 2 oxidation state increases.
- **Catenation** : The property of self linking of atoms of an element through covalent bonds to form straight or branched chains and rings of different size is called catenation. Among group 14 elements carbon shows maximum tendency for catenation. As we move down the group tendency to show catenation property decreases.
- **Formation of halides : (i) Tetrahalides** : All the elements of group 14 form tetrahalides of the general formula  $MX_4$  where X = F, Cl, Br or I. All these tetrahalides are covalent compounds and have tetrahedral structure. The stability of these tetrahalides decreases as we move from C to Pb *i.e.*,  $CX_4 > SiX_4 > GeX_4 > SnX_4 > PbX_4$   
(ii) **Dihalides** : All the elements of group 14 form dihalides of the formula  $MX_2$  where X = F, Cl, Br or I. The stability of these dihalides increases as we move down the group from C to Pb. Dihalides are generally ionic in nature and behave as reducing agents. The reducing character decreases in the order  $GeCl_2 > SnCl_2 > PbCl_2$ .
- **Formation of oxides** : All the elements of this group form two types of oxides (i) monoxides, (ii) dioxide.  
(i) **Monoxides** : All the elements of group 14 form monoxides of the general formula  $MO$  *i.e.*, C, SiO, GeO, SnO and PbO. These oxides except SiO and GeO are quite stable. CO is neutral while SnO and GeO are amphoteric.  
(ii) **Dioxides** : All these elements form dioxides of the general formula  $MO_2$ . All these dioxides can be prepared by heating the elements in oxygen except lead which forms lead monoxide.  $CO_2$  is a monomeric, linear molecule and hence exists as a gas while the dioxides of all other elements are crystalline solid with high melting points due to the formation of three dimensional network of bonds.
- **$p\pi-p\pi$  and  $p\pi-d\pi$  multiple bonding** : Amongst group 14 elements carbon shows a pronounced ability to form  $p\pi-p\pi$  multiple bonds with itself (*e.g.*, in graphite) and with other elements especially nitrogen and oxygen. Silicon and other heavier elements of this group show negligible tendency of this type.

$p\pi-p\pi$  multiple bonding has been observed particularly in case of silicon linked to oxygen and nitrogen. Thus trimethylamine  $\text{N}(\text{CH}_3)_3$  is pyramidal (N is  $sp^3$  hybridised) and is more basic whereas silicon compound trisilylamine  $\text{N}(\text{SiH}_3)_3$  is planar (N is  $sp^2$  hybridised) and is less basic. In this case, the lone pair of electrons on nitrogen atom is transferred to empty  $d$ -orbital of silicon  $p\pi-d\pi$  overlapping leading to planar structure.

- **Diamond** : In it C is  $sp^3$  hybridised.
- **Graphite** : In it C is  $sp^2$  hybridised.
- **Fullerenes** : In it C is  $sp^2$  hybridised.
- **Carbon monoxide** :  $2\text{C} + \text{O}_2 \xrightarrow{\Delta} 2\text{CO}$ . It is highly poisonous.
- **Carbon dioxide** :  $\text{C} + \text{O}_2 \xrightarrow{\Delta} \text{CO}$ . It causes green house effect.
- **Silicones** : Silicones are synthetic organosilican compounds containing repeated unit  $\text{R}_2\text{SiO}$  held by Si-O-Si linkages.

Silicones are water repellent, heat resistant, chemically inert, resistant to oxidation and attack by organic acids and are good electrical insulators.

## Group 13 Elements

### 1-Mark Questions

1. Mention two important ores of Boron.
2. Name the elements of group 13 which forms only covalent compounds.
3. Why the atomic radius of gallium is less than that of Al ?
4. Why does Boron forms electron deficient compounds ?
5. Boron does not exist as  $\text{B}^{3+}$  ion. Why ?
6. Why the trihalide of group 13 elements fume in moist air ?
7. Aluminium form  $[\text{AlF}_6]^{3-}$  but boron does not form  $[\text{BF}_6]^{3-}$ .
8. Why boric acid is a monobasic acid ?
9. White fumes appear around the bottle of anhydrous  $\text{AlCl}_3$ . Give reason.  
[NCERT]
10.  $\text{AlCl}_3$  exist as dimer while  $\text{BCl}_3$  exist as monomer, why ?
11. Mention the type of hybridization of Boron in  $\text{B}_2\text{H}_6$ . [Ans.  $sp^3$ ]
12. Write the formula of inorganic benzene.

13. Why aluminium utensils should not be kept in water overnight.[NCERT]  
14. Explain what happens when boric acid is heated.  
15.  $\text{BCl}_3$  exists but  $\text{BH}_3$  does not. Explain.

### Group 14 Elements

16. Why  $\text{SnCl}_4$  is more covalent than  $\text{SnCl}_2$  ?  
17. Why  $\text{PbCl}_4$  is good oxidising agent ?  
18. What are Germanes and Plumbanes ?  
19. Give one example of Zeolite.  
20. Mention the type of hybridization of carbon in diamond and graphite.  
21. Why  $\text{CCl}_4$  is insoluble in water but  $\text{SiCl}_4$  is soluble in water. Explain.  
22. Give two uses of silicones.  
23. Why graphite is used as lubricant ? [NCERT]  
24. Lead does not form  $\text{PbI}_4$ . Why ?  
25.  $\text{CO}_2$  is gas while  $\text{SiO}_2$  is solid at room temperature. Explain why ?  
26. Explain why silicon shows a higher covalency than carbon.  
27. Out of carbon and silicon which can form multiple bonds and why ?  
28. Write the formula of dry ice.  
29. What is the basic building unit of all silicates ?  
30. Graphite is a good conductor of electricity, but diamond is not. Why ?

### Group 13 Elements

#### 2-Mark Questions

1. Draw the structure of Diborane.  
2. What happens when :  
(a) Borax is heated strongly.  
(b) Boric acid is added to water.  
3. Write balanced chemical equations for :  
(a)  $\text{BF}_3 + \text{LiH} \longrightarrow$   
(b)  $\text{B}_2\text{H}_6 + \text{NH}_3 \longrightarrow$

4. Write chemical reactions to justify amphoteric nature of Al.
5. Suggest reason why the B-F bond length in  $\text{BF}_3$  and  $\text{BF}_4^-$  differ.
6. Give reason :
  - (i)  $\text{BF}_3$  act as weak Lewis acid.
  - (ii) Boron cannot show covalency more than four.
7. How can you explain higher stability of  $\text{BCl}_3$  as compared to  $\text{TlCl}_3$ .
8. Give reason :
  - (i) Aluminium alloys are used to make air craft body.
  - (ii) Aluminium wire is used to make transmission cables. [NCERT]
9. Describe the shapes of  $\text{BF}_3$  and  $\text{BH}_4^-$ . Assign the hybridization of boron in these species. [NCERT]
10. Explain the chemistry of borax bead test.

### Group 14 Elements

11.  $[\text{SiF}_6]^{2-}$  is known whereas  $[\text{SiCl}_6]^{2-}$  not. Give reason. [NCERT]
12. Hydrolysis of  $\text{SiCl}_4$  take place but of  $\text{CCl}_4$  does not. Why ?
13. Account for the following :
  - (a)  $\text{CO}_2$  is gas while  $\text{SiO}_2$  is solid at room temperature.
  - (b) Solid  $\text{CO}_2$  is known as dry ice.
14. Elemental silicon does not form graphite like structure as carbon does. Give reason.
15. Suggest a reason as to why CO is poisonous. [NCERT]
16. How is excessive content of  $\text{CO}_2$  responsible for global warming ? [NCERT]
17. What is allotropy ? Name two elements which exhibit allotropy.
18. Write equations for the production of water gas and producer gas from coke.
19. Define Zeolite. Name the zeolite which converts alcohols directly into gasoline.
- Arrange the hydrides of group 14 elements in increasing order of :
  - (a) Thermal stability
  - (b) Reducing power.

## Group 13 Elements

### 3-Mark Questions

1. Give reasons of the following :
  - (i) In diborane, two B—H—B bonds are different from common covalent bonds.
  - (ii) Aluminium metal shows amphoteric behaviour.
  - (iii) Quartz is used to develop extremely accurate clocks.
2. A certain salt X gives the following results :
  - (i) Its aqueous solution is alkaline to litmus.
  - (ii) It swells up to a glassy material Y on strong heating.
  - (iii) When conc.  $\text{H}_2\text{SO}_4$  is added to a hot solution of X, white crystal of an acid Z separates out. Write equations for all the above reactions and identify X, Y and Z.
3. Write balanced equation for :
  - (i)  $\text{B}_2\text{H}_6 + \text{H}_2\text{O} \longrightarrow$
  - (ii)  $\text{Al} + \text{NaOH} \longrightarrow$
  - (iii)  $\text{NaOH} + \text{B}_2\text{H}_6 \longrightarrow$
4. List two important properties in which Boron differs from the rest of the members of group. Mention the main reasons for the difference.
5. What are electron deficient compounds. Are  $\text{BCl}_3$  and  $\text{SiCl}_4$  electron deficient species ? Explain. [NCERT]

## Group 14 Elements

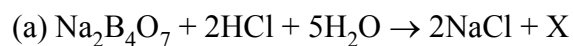
6. Select the member(s) of group 14 that :
  - (i) Forms the most acidic dioxide.
  - (ii) Is commonly found in + 2 oxidation state.
  - (iii) Used as semiconductor.

7. What are allotropes ? Sketch the structure of two allotropes of carbon namely diamond and graphite.
8. Give suitable reasons for the following :
- (a)  $\text{CO}_2$  turns lime water milky, but if passed for a long time, the solution becomes clear again.
  - (b) Graphite is a good conductor of electricity but diamond is insulator.
  - (c) Lead (IV) chloride is highly unstable towards heat.
9. (i) Write the resonance structure of  $\text{CO}_3^{2-}$  and  $\text{HCO}_3^-$ . [NCERT]  
(ii) Write the name of thermodynamically most suitable form of carbon.
10. (i) Explain why is there a phenomenal decrease in ionisation enthalpy from carbon to silicon ? [NCERT]  
(ii) Write an industrial application of silicones.

### Group 13 Elements

#### 5-Mark Questions

1. When metal X is treated with NaOH, a white precipitate 'A' is obtained, which is soluble in excess of NaOH to give soluble complex (B). Compound 'A' is soluble in dilute HCl to form compound 'C'. The compound 'A' when heated strongly gives 'D', which is used to extract metal. Identify X, A, B, C and D. Write suitable equations to support their identities. [NCERT]
2. (i) If B-Cl bond has dipole moment explain why  $\text{BCl}_3$  molecule has zero dipole moment.  
(ii) A mixture of dil. NaOH and aluminium pieces is used to open drain. Give reason.  
(iii) Aluminium wire is used to make transmission cables. Why ? [NCERT]
3. (i) Identify the compounds X and Y in the following reactions :



(ii) Write the name of group 13 element which is used to measure high temperature.

(iii) Why in case of Thallium + 1 oxidation state is more stable than + 3.

### Group 14 Elements

4. Compare the general trend in the following properties of the elements of group 13 and 14 :

(a) Atomic size, (b) Ionisation enthalpy, (c) Metallic character, (d) Oxidation states, (e) Nature of halides.

5. Name the following :

(a) The crystalline form of silica used in modern radio and T.V. broadcasting and mobile radio communication.

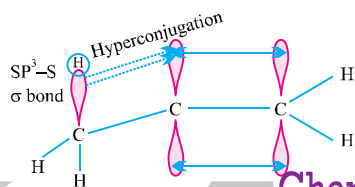
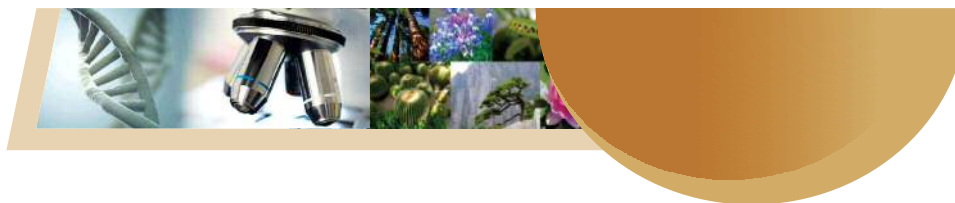
(b) The oxides of carbon which form a complex with haemoglobin 300 times more faster than oxygen.

(c) The allotrope of carbon which has  $\text{O}_f\text{H}^\ominus = 0$ .

(d) A type of polymer is semiorganic in nature.

(e) Two man made silicates.

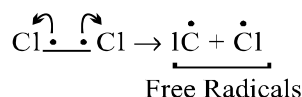




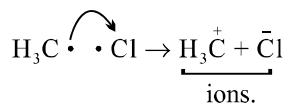
## Chapter - 12

# Organic Chemistry: Some Basic Principles And Techniques

- The branch of chemistry which deals with hydrocarbons and their derivatives is called organic chemistry.
- Carbon forms large number of organic compound because of its properties of catention and tetravalency.
- **Functional Group** : An atom or a group of atoms joined in a specific manner, which provides certain characteristics chemical properties to the organic compounds, is called functional group.
- **Homologues** : A group or series of an organic compound each containing a characteristics functional group from a homologous series and the members of the series are called “homologues”.
- **Fission of a Covalent Bond** :
  - (a) **Homolytic cleavage** : In this cleavaage, one of the electrons of the shared pair in a covalent bond goes with each of the bonded atoms.

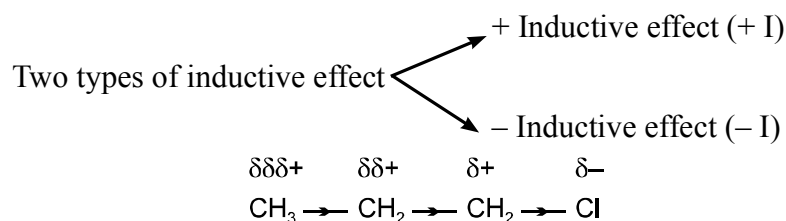


- (b) **Heterolytic cleavage** : In heterolytic cleavage the bond breaks in such a fashion that the shared pair of electrons remains with one of the fragment.

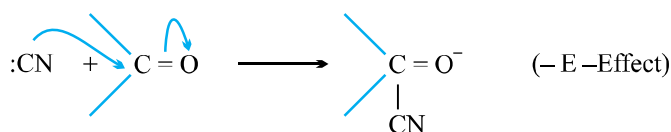
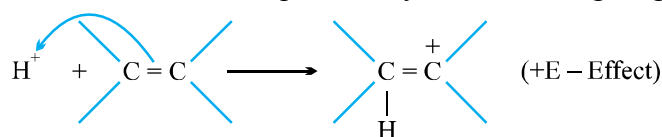


- **Electron Displacement Effects in Bivalent Bonds.**

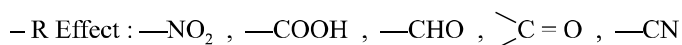
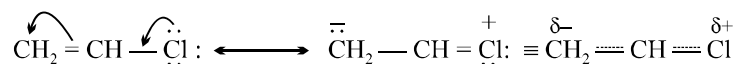
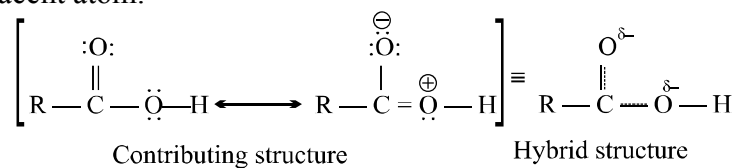
**Inductive effect (I)** : Polarisation of a bond caused by the polarisation of adjacent bond is referred to as the inductive effect.



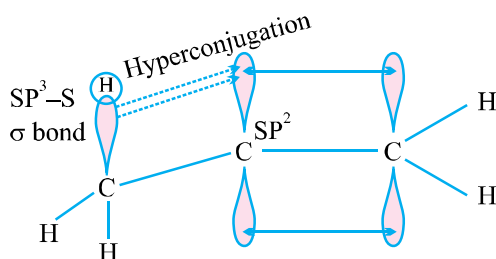
- It is a permanent effect and decrease with the increase in distance.
- I effect :  $-\text{NO}_2 > -\text{F} > -\text{Cl} > -\text{Br} > -\text{I} > -\text{OCH}_3 > -\text{C}_6\text{H}_5$
- + I effect :  $-\text{C}(\text{CH}_3)_3 > -\text{CH}(\text{CH}_3)_2 > -\text{C}_2\text{H}_5 > -\text{CH}_3$
- Electromeric effect** : The complete transfer of the shared pair of  $\pi$  electrons of a multiple bond to one of the atoms in the presence of the attacking reagent is called electromeric effect.
- If the transference of  $e^-$  towards attacking reagent + **E effect**.
- If the transference of  $e^-$  takes place away from attacking reagent - **E effect**.



- Resonance effect (+ R effect)** : The polarity produced in the molecule by the interaction of two  $\pi$ -bond and lone pair of electrons present on an adjacent atom.



- Hyper conjugation** : It is special kind of resonance in which delocalisation of  $e^-$  takes place through overlap between  $\sigma$ -bond orbital and  $\pi$ -orbital. It is also called no bond resonance.

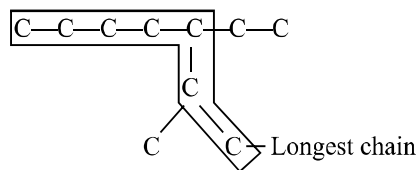
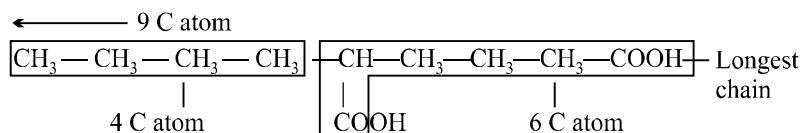


- **IUPAC Nomenclature of Organic Compounds :** Following rules are used to write the IUPAC name of an organic compound.

**Rule 1.**

- **Longest chain rule :** The chain containing the principal functional group, secondary functional group and multiple bonds as many as possible is the longest possible chain.

In the absence of functional group, secondary group and multiple bonds, the chain containing the maximum number of C-atoms will be the longest possible chain *e.g.*,



Choose the word root from the table given below for the longest possible chain.

**Word Root for Carbon Chain**

Chain length	Word root	Chain length	Word root
C <sub>1</sub>	Meth-	C <sub>7</sub>	Hept
C <sub>2</sub>	Eth-	C <sub>8</sub>	Oct
C <sub>3</sub>	Prop-	C <sub>9</sub>	Non
C <sub>4</sub>	But-	C <sub>10</sub>	Dec
C <sub>5</sub>	Pent-	C <sub>11</sub>	Undec
C <sub>6</sub>	Hex-	C <sub>12</sub>	Dodec

### Rule 2

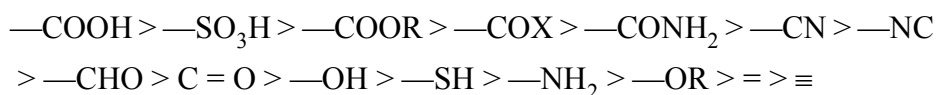
- **Lowest number rule :** Numbering is done in such a way so that :

(1) branching if present gets the lowest number.

(2) the sum of numbers of side chain is lowest.

(3) principal functional group gets the lowest number.

Select the principal functional group from the preference series :



Functional group other than the principal functional group are called substituents.

### Rule 3

- **Naming the prefixes and suffixes :** Prefix represents the substituent and suffix is used for principal functional group.

Primary prefixes are cyclo, bicyclo, di, tri, tetra, tetrakis etc.

### Secondary prefixes are tabulated below :

Substituent	Prefix	Substituent	Prefix
—F	Fluoro	—N=N—	diazo
—Cl	Chloro	—N=O	nitroso
—Br	Bromo	—NO <sub>2</sub>	nitro

Primary suffix are **ene**, **ane** or **yne** used for double, single and triple bonds respectively.

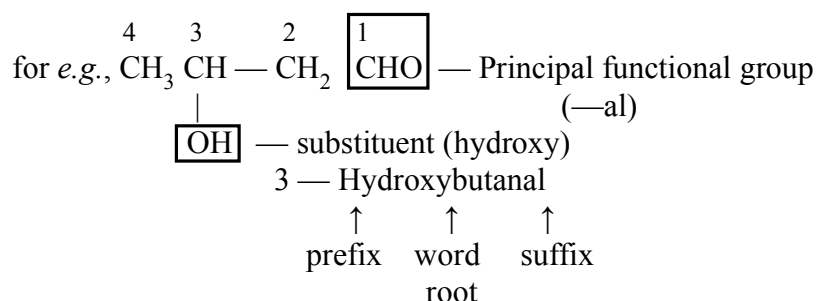
### Secondary suffixes are tabulated below :

S.No.	Class	Formula	Prefix	Suffix
1.	Acid halides	$\begin{array}{c} \text{O} \\    \\ \text{—C—X} \end{array}$	halocarbonyl	—oyl halide
2.	Alcohols	—OH	hydroxy	—ol
3.	Aldehydes	—CHO	formyl —carbaldehyde	—al
4.	Ketones	C = O	oxo	—one
5.	Amides	—CONH <sub>2</sub>	carbamoyl	—amide

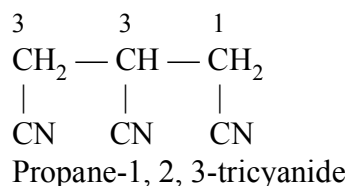
6.	Amine	$\text{—NH}_2$	amino	$\text{—amine}$
7.	Carboxylic acid	$\text{—COOH}$	carboxy	$\text{—carboxylic acid}$
8.	Ester	$\text{—COOR}$	alkoxy carbonyl	$\text{—alkyl alkanoate}$
9.	Nitriles	$\text{—CN}$	cyano	$\text{—nitrile}$
10.	Sulphonic acid	$\text{—SO}_2\text{—OH}$	sulpho	$\text{—sulphonic acid}$

Here according to the rules, given above, the IUPAC name of a compound can be written as  $\Rightarrow$  Prefixes + Word root + Suffixes.

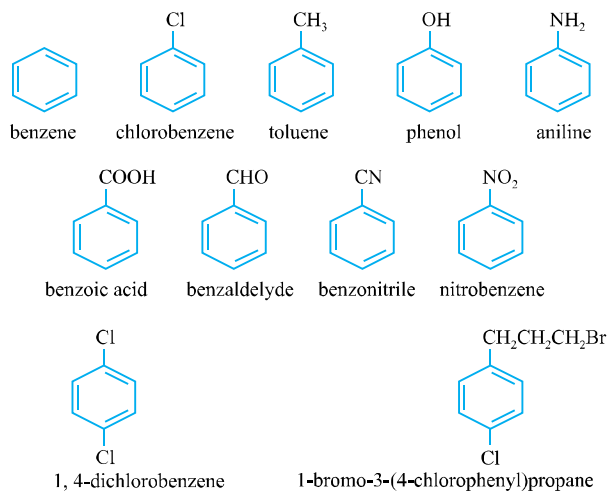
- **Primary prefix + secondary prefix + Word root + primary suffix + secondary suffix**



- If more than two similar functional groups are present, all the groups are considered as substituent, for *e.g.*,



- **Naming of Aromatic compounds :** IUPAC accepted their common trivial names *e.g.*,

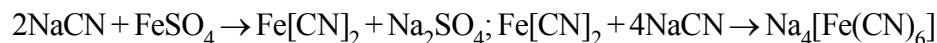


- **Purification and Characterization of Organic Compounds :**

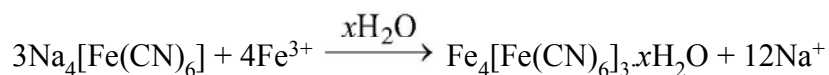
**(1) Lassaigne's test for nitrogen :** Lassaigne's extract is heated with  $\text{FeSO}_4$  solution in presence of alkali, the solution is cooled and acidified with dil.  $\text{H}_2\text{SO}_4$ . If a green or blue colouration is obtained, it confirms the presence of N in the organic compound. The chemistry of the test is :



From organic compound



Sod. ferrocyanide



Ferric ferrocyanide (Prussian Blue)

This test is very delicate and is given by all compounds containing C and N.  $\text{NH}_2\text{NH}_2$ ,  $\text{NH}_4\text{C}$ ,  $\text{NaNO}_3$  etc. do not respond to this test since they do not contain carbon.

Formation of blood red colour indicates the presence of both N and S.



From organic compound

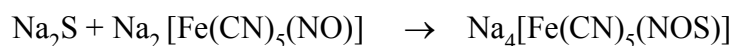
Sod. thiocyanate or Sod. sulphocyanide



Ferric thiocyanate (blood red colour)

- **Detection of sulphur :** If S is present, during fusion with Na metal,  $\text{Na}_2\text{S}$  is formed which may be tested as follows :

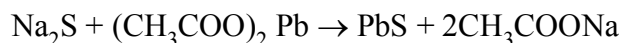
(i) With sodium nitroprusside violet colouration is produced :



Sodium nitroprusside

(Violet colour)

(ii) With lead acetate, black ppt. of  $\text{PbS}$  is formed.



- **Detection of halogens :**

**(a) Lassaigne's test :** When the organic compounds is fused with Na metal, the halogens combine with Na to form sodium halides.

The presence of these halides is tested with  $\text{AgNO}_3$  solution.

(i) A white ppt. soluble in  $\text{NH}_4\text{OH}$  indicates chlorine.

(ii) A pale yellow ppt. partially soluble in ammonia indicates bromine.

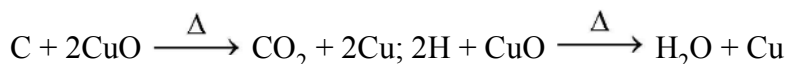
(iii) A yellow ppt. insoluble in ammonia indicates iodine.

If the organic compound also contain N or S, the sodium extract is first boiled with dil.  $\text{HNO}_3$  to decompose any cyanides or sulphides, otherwise these will form ppt. with  $\text{AgNO}_3$  solution.

- **Detection of phosphorus :** Phosphorus is detected by fusing the organic compound with sodium peroxide, in which phosphorus is converted into sodium phosphate.

The fused mass is extracted with  $\text{H}_2\text{O}$  and then boiled with conc.  $\text{HNO}_3$  and then ammonium molybdate is added. Appearance of yellow ppt. or yellow colouration due to the formation of ammonium phosphomolybdate indicates the presence of phosphorus.

- **Estimation of carbon and hydrogen : Liebig's method :** A known mass of the organic compound is heated strongly with excess of dry copper oxide in a current of dry air or oxygen (free from  $\text{CO}_2$ ) when carbon present in the organic compound compound is oxidised to  $\text{CO}_2$  and hydrogen to  $\text{H}_2\text{O}$ .

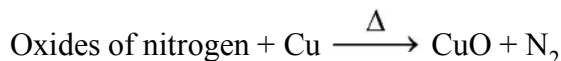
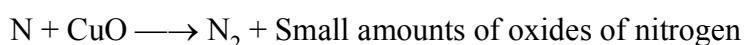
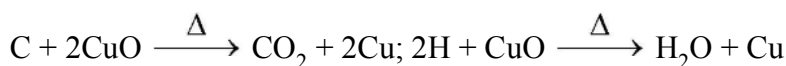


$$\text{Percentage of carbon} = \frac{12}{44} \times \frac{\text{Mass of CO}_2 \text{ formed}}{\text{Mass of substance taken}} \times 100$$

$$\text{Percentage of carbon} = \frac{2}{18} \times \frac{\text{Mass of H}_2\text{O formed}}{\text{Mass of substance taken}} \times 100$$

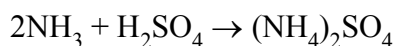
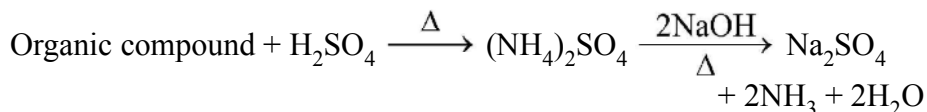
- **Estimation of nitrogen :**

(a) **Dumas method :**



$$\text{Percentage of nitrogen} = \frac{28}{22400} \times \frac{\text{Vol. of N}_2(\text{in cm}^3) \text{ at STP}}{\text{Mass of substance taken}} \times 100$$

(b) **Kjeldahl's method :**



**Percentage of nitrogen :**

$$= \frac{1.4 \times \text{Molarity of the acid} \times \text{V ol. of acid used} \times \text{Basicity of the acid}}{\text{Mass of substance taken}}$$

● **Estimation of halogens : (Carius method) :**

$$\text{Percentage of chlorine} = \frac{35.5}{14.5} \times \frac{\text{Mass of AgCl formed}}{\text{Mass of substance taken}} \times 100$$

$$\text{Percentage of bromine} = \frac{80}{188} \times \frac{\text{Mass of AgBr formed}}{\text{Mass of substance taken}} \times 100$$

$$\text{Percentage of iodine} = \frac{127}{235} \times \frac{\text{Mass of AgI formed}}{\text{Mass of substance taken}} \times 100$$

● **Percentage of sulphur** =  $\frac{32}{233} \times \frac{\text{Mass of BaSO}_4 \text{ formed}}{\text{Mass of substance taken}} \times 100$

- **Estimation of phosphorus :** A known mass of the organic compound is heated with fuming  $\text{HNO}_3$  in a Carius tube when phosphorus of the organic compound is oxidized to  $\text{H}_3\text{PO}_4$ . Phosphoric acid thus formed is precipitated as magnesium ammonium phosphate by adding magnesia mixture (a solution containing  $\text{MgCl}_2$ ,  $\text{NH}_4\text{Cl}$  and  $\text{NH}_4\text{OH}$ .)

$$\text{Percentage of phosphorus} : \frac{62}{222} \times \frac{\text{Mass of Mg}_2\text{P}_2\text{O}_7 \text{ formed}}{\text{Mass of substance taken}} \times 100$$

- **Estimation of oxygen :** A definite mass of an organic compound is decomposed by heating with  $\text{N}_2$  gas. The mixture is then passed over red hot coke when all oxygen is converted to  $\text{CO}$ . This mixture is then passed through  $\text{I}_2\text{O}_5$  when  $\text{CO}$  is oxidised to  $\text{CO}_2$  producing iodine. The % of oxygen can be derived from the amount of  $\text{CO}_2$  or  $\text{I}_2$  produced.

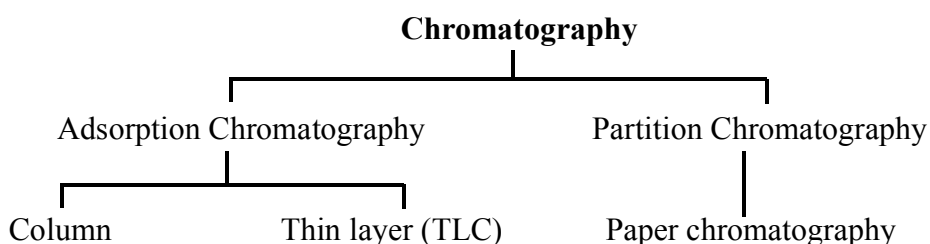
$$\text{Percentage of oxygen} : \frac{16}{44} \times \frac{\text{Mass of CO}_2 \text{ formed}}{\text{Mass of substance taken}} \times 100$$

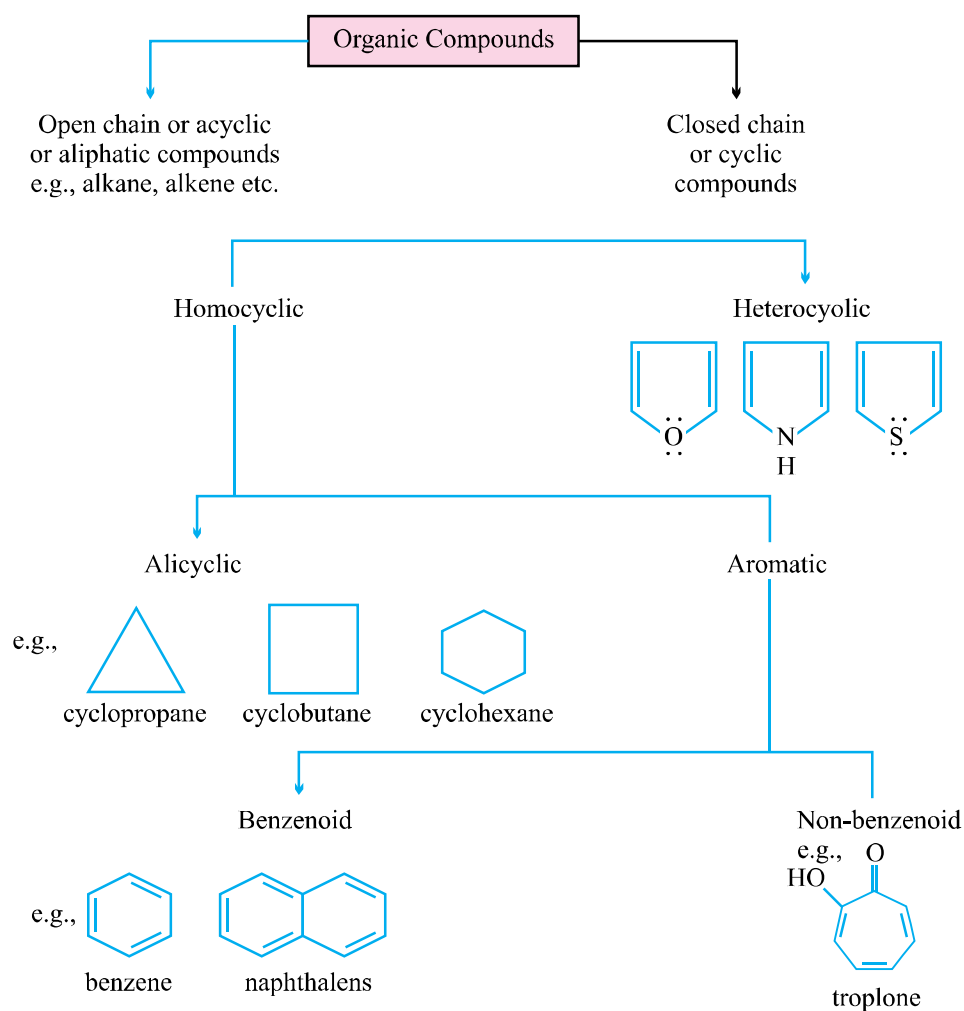
**Method of Purification of organic compounds :**

- **Crystallisation :** Process of solidification of a pure substance from its dissolved state. This method is based upon differences in their solubility in a given solvent or in mixture of solvents.



- **Sublimation** : It is a process of conversion of a solid into gaseous state on heating without interchanging into liquid. The process is used for the separation of volatile solids, which sublime on heating from the non-volatile solids.
- **Distillation** : It is a process of conversion of a liquid into vapours by heating followed by condensation of vapours. The method is used for the purification of liquids which boil without decomposition and are present with non-volatile impurities.
- **Fractional distillation** : Process used to separate mixture of two or more miscible liquids having different boiling points. It is mainly used in distillation of petroleum, coaltar and crude alcohol.
- **Distillation under reduced pressure** : This process is used when the liquid has a tendency to decompose near its boiling point. Under reduced pressure, the liquid will boil at a low temperature without decomposing.
- **Steam distillation** : Purification of a substance from non-volatile impurities provided the substance itself is volatile in steam and insoluble in water.
- **Chromatography** : Technique of separating the constituents of a mixture by the differential movement of individual components through the stationary phase under the influence of mobile phase. Two types of chromatography.





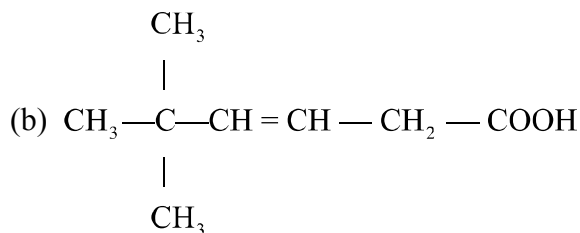
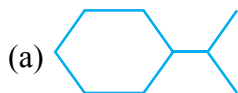
## 2-Mark Questions

6. How many  $\sigma$  and  $\pi$  bonds are present in each of the following molecules ?  
(a)  $\text{HC} \equiv \text{CCH} = \text{CHCH}_3$  (b)  $\text{CH}_2 = \text{C} = \text{CH}-\text{CH}_2\text{CH}_3$   
[Ans. (a)  $\sigma = 10$ ,  $\pi = 3$  (b)  $\sigma = 12$ ,  $\pi = 2$ ]
7. Mention the hybridisation of  $\text{C}^*$  and shape of the compound.  
(a)  $\text{H}_2\text{C}^* = \text{O}$  (b)  $\text{CH}_3-\text{C}^* \equiv \text{N}$
8. Which bond is more polar in the following pair of molecules :  
(a)  $\text{CH}_3-\text{H}$  or  $\text{CH}_3-\text{Br}$  (b)  $\text{CH}_3-\text{NH}_3$  or  $\text{CH}_3-\text{OH}$
9. Draw formula of first four members of homologous series beginning with the compound  $\text{CH}_2 = \text{CH}_2$ .
10. (a) Why does carbon exhibit catenation to maximum extent ?  
(b) Give hybridization of each carbon in following compound  $\text{CH}_2 = \text{CH}-\text{CN}$ .

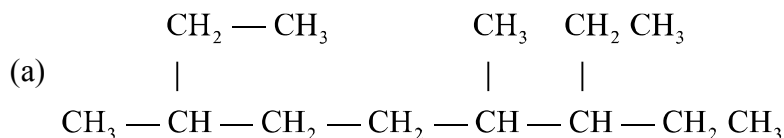
## 1-Mark Questions

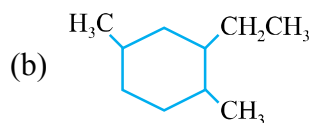
### Nomenclature

1. Write the IUPAC name of

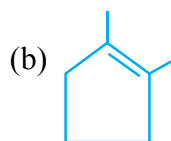


2. Draw structure of 3-isopropyl-2-methylhexane.  
3. Write the structure of compound that contains both  $1^\circ$  and  $2^\circ$  alcohol.  
4. Give IUPAC name of following :



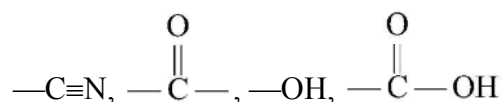


5. Give IUPAC name of following bond-line formulae :



### 2-Mark Questions

6. Write the correct order of priority of the following functional groups :

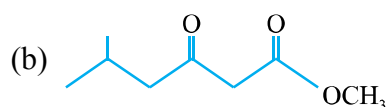
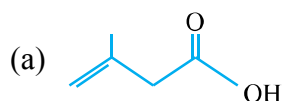


7. Write the structural formula of :

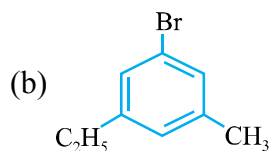
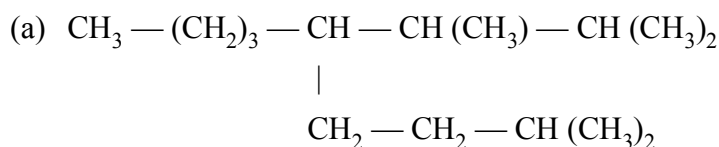
(i) O-Ethylanisole

(b) 4-ethyl-1-fluoro-2-nitrobenzene

8. Identify the functional groups in :



9. Give IUPAC name of following :



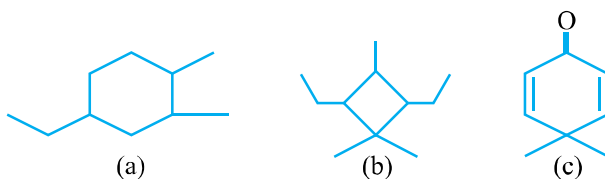
10. Give condensed and bond-line structural formulae and identify the functional group(s) present, if any for :

(a) Cyclo octa-1, 5-diene

(b) 2(4-isobutylphenyl) propanoic acid

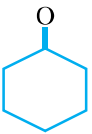
### 3-Mark Questions

11. Draw the structure of :
- Pent-3-enoic acid
  - 4-Methylpentanone
  - 4-Ethyl-3-fluorophenol.
12. Write the IUPAC name of the following compound :



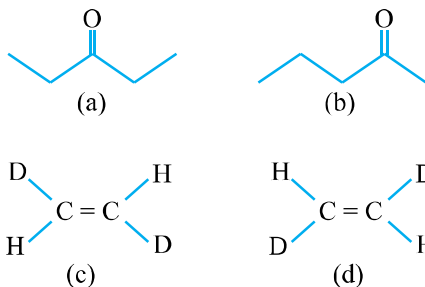
### 1 Mark Questions

### Isomerism

- Write functional isomer of molecular formula  $C_3H_6O$ .
- Write tautomeric form of following structures :
  - $CH_3 - \overset{\overset{O}{\parallel}}{C} - CH_3$
  - 
- Identify the chiral carbon in the given compound  $CH_3 - CHOH - CH_2 - CH_3$ .

### 2 Mark Questions

4. What is the relationship between the members of the following pairs of structures ?



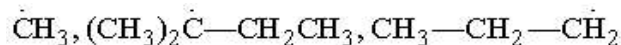
5. Write all the possible isomers of the aromatic compound  $C_8H_{10}$ .

### 1 Mark Questions

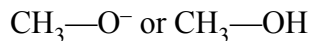
### Concepts in Reaction Mechanism

- Identify electrophilic centre in  $CH_3CHO$ .
- Identify nucleophilic centre in  $CH_3Br$ .

3. Arrange the following in decreasing order of stability :



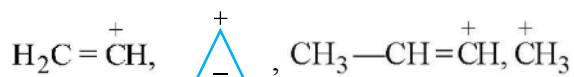
4. Which species can act as an acid and why ?



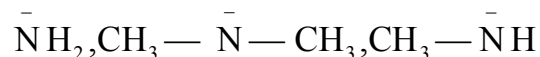
5. What types of attacking reagents are produced by heterolytic cleavage of bonds ?

6. Out of  $\text{CH}_3\text{COOH}$  and  $\text{NO}_2\text{CH}_2\text{COOH}$  which is more acidic in nature and why ?

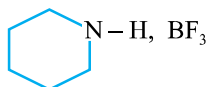
7. Identify the most stable carbonation among the following :



8. Identify the weakest nucleophile among the following :



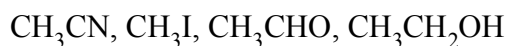
9. Select the nucleophile and electrophile in the following :



10. Give reason  $(\text{CH}_3)_3\text{C}^+$  is more stable than  $\text{CH}_3\overset{+}{\text{C}}\text{H}_3$  and  $\overset{+}{\text{C}}\text{H}_3$ .

### 2 Mark Questions

11. Mark the electrophilic centre in the following molecules :

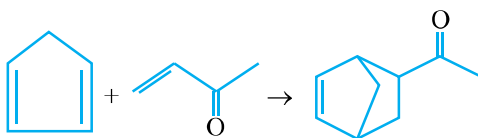


12. Benzyllic free radical is more stable than allylic free radical. Explain with resonance.

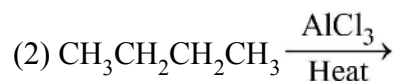
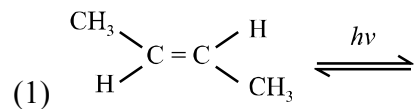
13. Classify each of the following carbon intermediates :



14. Classify whether the following reaction is rearrangement addition, or elimination ?



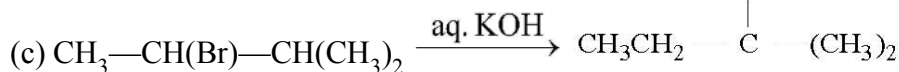
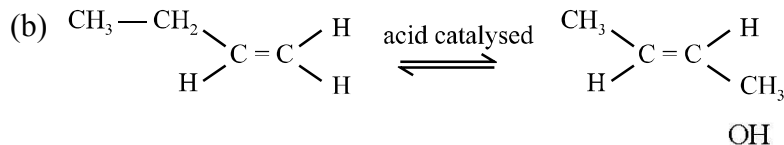
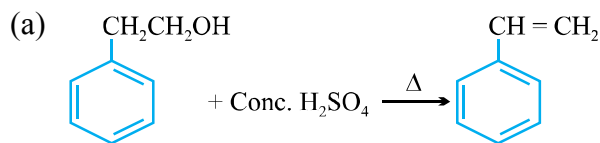
15. Write the product of following reaction.



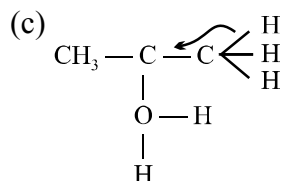
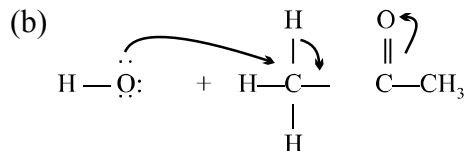
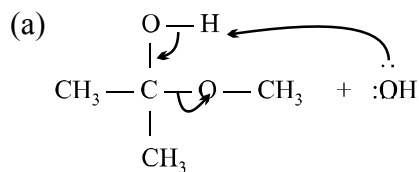
### 5 Mark Questions

16. Write structure of various carbocation that can be obtained from 2-methylbutane. Arrange these carbocation in order of increasing stability.

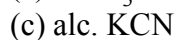
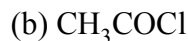
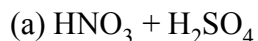
17. Classify the reaction type as elimination, rearrangement addition and substitution.



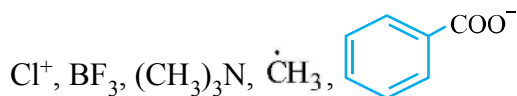
18. Follow the flow of electrons indicated by the curved arrows and predict the products :



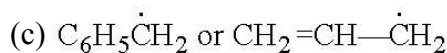
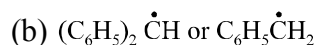
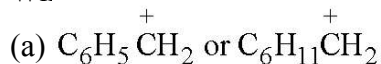
19. Name the electrophile/nucleophile generated by following species :



20. Identify the nucleophiles, electrophiles and free radicals amongst the following :

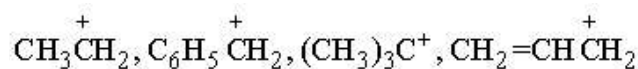


21. Which is more stable and why ?



### 5 Mark Questions

1. Arrange the following according to given property :



(decreasing order of stability)


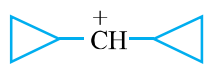
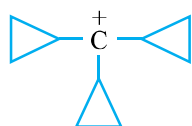
2.  $\text{HC}\equiv\bar{\text{C}}$ ,  $\text{CH}_2=\bar{\text{C}}\text{H}$ ,  $\text{CH}_3-\bar{\text{C}}\text{H}_2$ ,  $\bar{\text{C}}\text{H}_3$  (increasing order of stability)

3.  $\text{C}_6\text{H}_5\dot{\text{C}}\text{HCH}_3$ ,  $\text{C}_6\text{H}_5\text{CH}_2\dot{\text{C}}\text{H}_2$ ,  $\text{C}_6\text{H}_5\dot{\text{C}}(\text{CH}_3)_2$

(increasing order of stability)

4.  $\text{CH}_3\text{CH}^+-\text{CH}_3$ ,  $\text{CH}_3-\text{CH}^+-\text{OCH}_3$ ,  $\text{CH}_3\text{CH}^+-\overset{\text{O}}{\parallel}{\text{C}}-\text{OCH}_3$

(decreasing order of stability)

5.    (decreasing order of stability)

### 2-Mark Questions

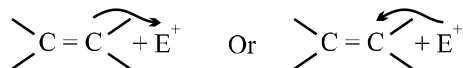
#### Electronic Displacement in Covalent Bond

1. Name the kind of effect that operates to explain the stability of carbocations.
2. Why inductive effect is also called transmission effect ?

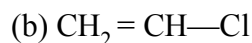
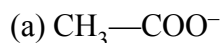


3. Which permanent effect of organic compound is also known as 'No bond resonance effect' ?

4. Which is correct and why ?

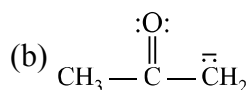
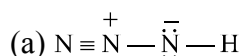


5. Write resonating structure of the following and show the movement of electron by curved arrows :

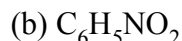


### 2 Mark Questions

6. Draw the resonating structure of :

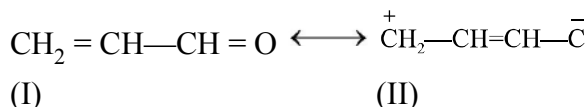


7. Write resonance structure of :



8. Explain why alkyl groups acts as  $e^-$  donar when attached to a  $\pi$ -system.

9. Resonance structures of propenal are given below. Which of these resonating structure is more stable ? Give reason.



10. Explain the following terms :

(a) Electromeric effect

(b) Hyperconjugation

11. (a) Explain + I and – I effect.

(b) Select the group giving + I effect and – I effect from the following list :

(i)  $-\text{NO}_2$  (ii)  $-\text{CN}$  (iii)  $\text{Cl}^-$  (d)  $\text{CH}_3^-$

12. Explain the importance of inductive effect in determination of acidic or basic strength of substances.

13. Write resonance structure of  $\text{CH}_2=\text{CH—CHO}$ . Indicate relative stability of the contributing structures.

### 5-Mark Questions

1. Give reason for the following :

(a) Chlorobenzene is  $o^-$  and  $p^-$  directing towards the electrophillic substitution reaction.

(b) Inductive effect decrease with the increase in distance.

(c) Hyperconjugation effect is extended form of resonance effect.

2. Arrange the following according to given property :
- (a)  $\text{—NO}_2$ ,  $\text{—COOH}$ ,  $\text{—F}$ ,  $\text{—CN}$ ,  $\text{—I}$  (increasing order of  $-I$  effect)
- (b)  $\text{CH}_3\text{—}$ ,  $(\text{CH}_3)_2\text{C—}$ ,  $(\text{CH}_3)_2\text{CH—}$ ,  $\text{CH}_3\text{CH}_2\text{—}$   
(decreasing order of  $+I$  effect)

### Purification Methods

#### 1 Mark Questions

1. What conditions must be satisfied by a suitable solvent in the crystallization method ?
2. Which technique can be used for purification of iodine that contains traces of NaCl ?
3. When do we use hot water funnel for filtration ?  
[Hint : When organic substance crystallises during filtration.]
4. A liquid (10 mL) has three components A, B, C. Which technique is most suitable to separate A, B, C from such a small amount of mixture ?
5. A substance has boiling point 355 K, but it starts decomposing near this temp. Which type of distillation process is suitable for its purification ?
6. Name the adsorbent used in column chromatography.
7. Which technique can be used to separate naphthalene from kerosene oil present in its mixture ?
8. A mixture contains nitrobenzene and benzoic acid. How can this mixture be separated into its constituents by technique of extraction using appropriate chemical reagent ?

[Hint : By using hot water as solvent and adopting differential extraction.]

#### 2 Mark Questions

9. The  $R_f$  value of A and B in a mixture determined by TLC in a solvent mixture are 0.65 and 0.42 respectively. If the mixture is separated by column chromatography using the same solvent mixture as a mobile phase, which of the two components A or B will elute first ? Explain.  
[Hint : A will elute first because it has more  $R_f$  value.]
10. Name a suitable technique of separation of the components from a mixture of :
  - (a) Water and aniline.
  - (b) Methanol and Propanone.

## Qualitative and Quantitative Analysis

### 2-Mark Questions

1. A student was given the compound  $\text{C}_6\text{H}_4(\text{NH}_2)\text{SO}_3\text{H}$  for elemental analysis, while performing Lassaigne's test for N, what colour will he get and why ?
2. Why diazonium salts do not show sodalime test for nitrogen ?  
[Hint : Because salts do not liberate  $\text{NH}_3$  gas under there conditions.]
3. What is the function of fusing the organic compound with sodium metal ?
4. If silver nitrate solution is added to chlorobenzene, will there be formation of white ppt.
5. Name the oxidising agent used in the combustion tube along with the organic compound.  
[Ans. CuO]

### 2-Mark Questions

6. Why is it necessary to boil Lassaigne's extract with  $\text{HNO}_3$  before testing it for halogens ?
7. 0.25 g of an organic compound containing C, H and O was analysed by the combustion method. The increase in the man of calcium chloride tube and the potash bulbs at the end of the operation was found to be 0.15 g and 0.1837 g respectively. Calculate the percentage composition of the compound.
8. Will  $\text{CCl}_4$  give white precipitate of  $\text{AgCl}$  on heating it will silver nitrate ? Give reason.
9. For which type of compounds Kjeldahl's method is not applicable ?
10. 0.90 of an organic compound on combustion 2.64 g of  $\text{CO}_2$  and 0.63 g of  $\text{H}_2\text{O}$ . Calculate the percentage of C and H in the compound.

### 3-Mark Questions

11. What will happens if a student acidifies the Lassaigne's extract with dil.  $\text{H}_2\text{SO}_4$  in place of dilute  $\text{HNO}_3$ . Write the reaction involved.
12. (a) In DNA and RNA, nitrogen atom is present in the ring system. Can Kjeldahl's method is used for the estimation of N-present in these. Give reasons.  
(b) Why is it necessary to use ethanoic acid and not sulphuric acid for

acidification of sodium extract for testing sulphur by lead acetate test ?

13. 0.2325 g of an organic compound was analysed for nitrogen by Duma's method. 0.0317 L of moist nitrogen was collected at 25°C and 755.8 mm pressure calculate the percentate of nitrogen.

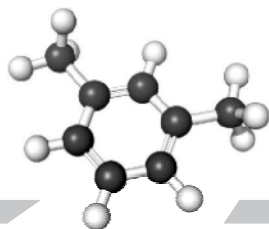
[Ans. Aq. tension of water at 25°C is 23.8 mm Hg]

### 5-Mark Questions

14. (a) Out of the different gases formed in Duma's method, which gas is not observed over an aqueous solution of KOH.
- (b) What is the function of adding small amount of  $K_2SO_4$  and a little amount of Hg or  $CuSO_4$  is Kjeldahl's flask ?
- (c) Explain why a solution of KOH used to absorb  $CO_2$  evolved during the estimation of carbon in an organic compound.
- (d) An organic compound contain diazo group ( $-N=N-$ ) or nitro group or 'N' in the ring. Name the method used to estimate nitrogen in the compound.
- (e) In victor Meyer's method, what is the gas collected in the gas jar ?

[Hint : Oxygen gas]

15. (a) 0.4 g of the compound was Kjeldahl's and ammonia evolved was absorbed into 50 ml of  $\frac{M}{4}$   $H_2SO_4$  solution. The residual acid solution was diluted with distilled water and the volume was made upto 150 ml. 20 ml of this diluted solution required 31 ml of  $\frac{M}{20}$  NaOH solution for complete neutratization. Calculate the % of N is compound.[Ans.46.8%]
- (b) Write the formula for the prursian blue colour obtained during Lassaigne's test for nitrogen.
- (c) Give test to detect the presence of sulphur in compound.

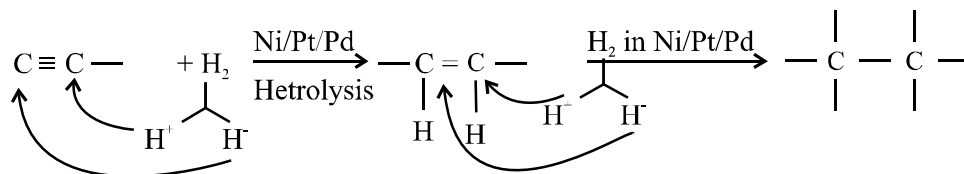
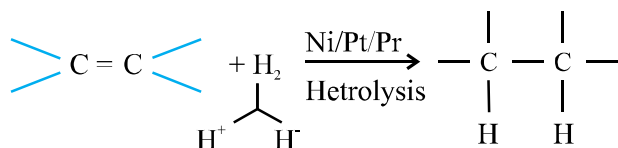


## Chapter - 13

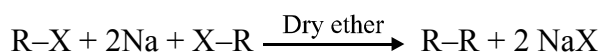
# Hydro Carbons

### Preparation of Alkanes :

(1) From unsaturated hydrocarbons :

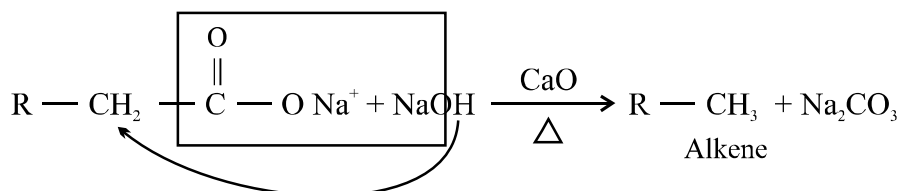
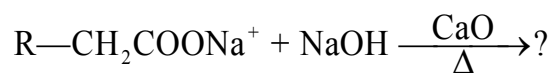


(2) Wurtz reaction  $\left[ \text{Alkyl Halide} + \text{Na} \xrightarrow[\text{ether}]{\text{Dry}} \text{Higher alkane} \right]$

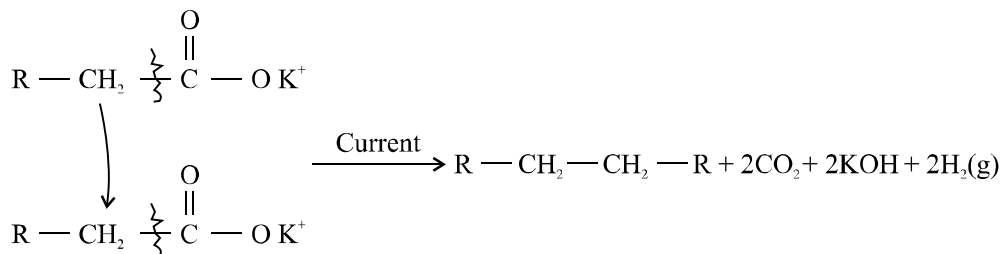
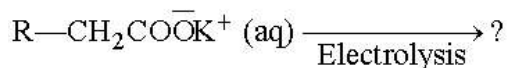
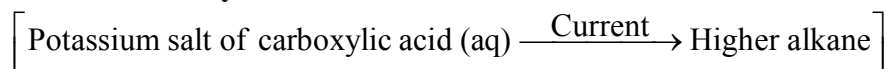


(3) Sodalime decarboxylation method :

[Sodium salt of carboxylic acid + NaOH + CaO  $\rightarrow$  Alkane [ $n_c = 1$  less]]

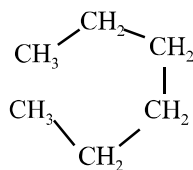
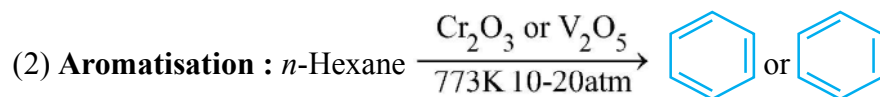
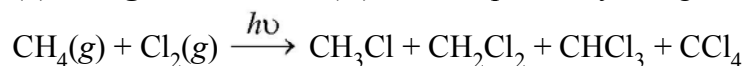


● **Koble's electrolytic method :**



● **Chemical Properties of Alkanes :**

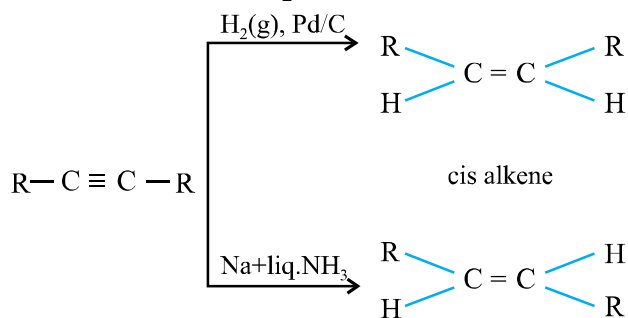
(1) **Halogenation :** One (H) atm is replaced by halogen at a time.



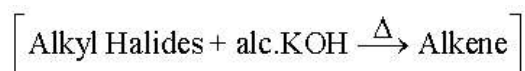
## Alkenes

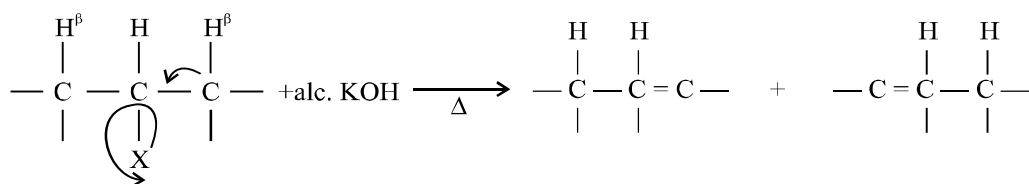
● **Preparation of alkenes :**

(1) From alkynes [Alkyne + H<sub>2</sub> → Alkene]



(2) From alkyl halide by (dehydrohalogenation)



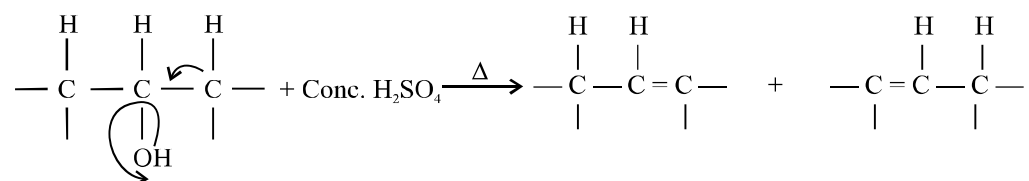


Carbon attached with halogen is  $\alpha$ -carbons

Carbon attached with  $\alpha$ -carbons is  $\beta$ -carbons

Halogen is removed and 'H'-atom is removed from  $\beta$ -carbon to form (C = C) double bond.

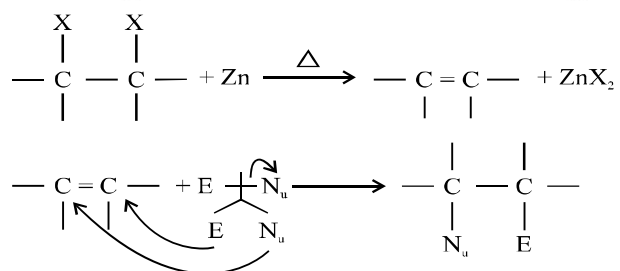
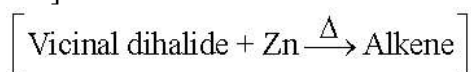
(3) By Dehydration of alcohols (Ion of water molecule)



Carbon attached to alcoholic group is  $\alpha$ -carbon.

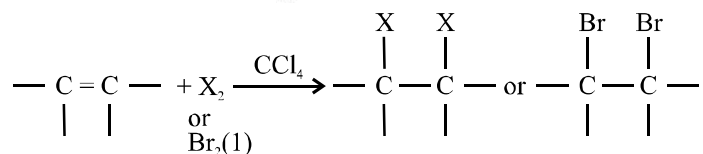
Carbon attached to  $\alpha$ -carbon is  $\beta$ -carbon.

(4) From vicinal dihalides [Compounds in which halogen atom are attached with adjacent carbons]



### Chemical Properties of Alkenes :

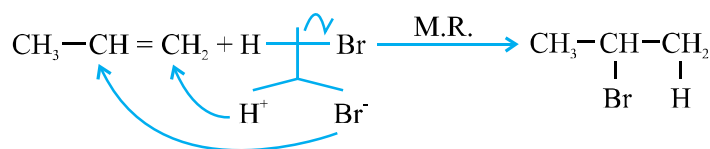
(1) Addition of Halogens :  $[\text{Alkene} + \text{X}_2 \longrightarrow \text{Vicinal dihalide}]$



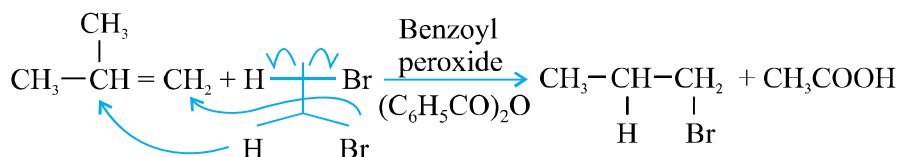
(2) Addition of H—X :  $[\text{Alkene} + \text{HX} \longrightarrow \text{Alkyl halide}]$

(A) **Markownikoff's rule (M.R.)** : During electrophilic addition of hydrogen halide, the electron deficient electrophile ( $\text{E}^+$ ) always attack

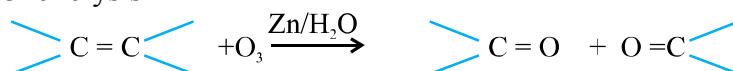
on that doubly/triply bounded carbon atom. which already has greater number of hydrogen atoms.



(B) **Peroxide/Kharasch effect (Anti M.Rule)** : This effect takes place in presence of peroxides when the hydrogen free radical ( $\dot{\text{H}}$ ) attacks on that doubly bonded carbon which has lesser number of hydrogen atoms.



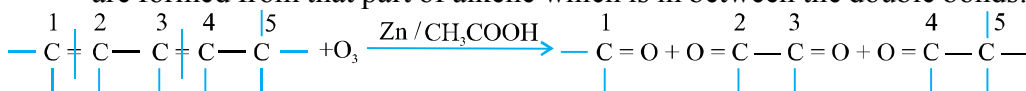
(C) **Ozonolysis**



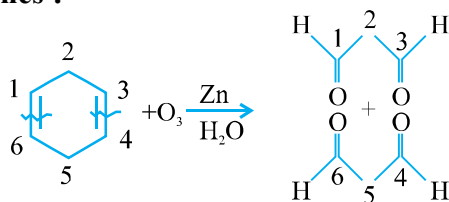
In this reaction all those carbons which form double bonds get findly converted into carbonyl carbons.

If alkenes are symmetrical then both carbonyl compounds are same.

If more than two double bonds are present then we get atleast one compound which has two carbonyl groups at the end. Such bifunctional compounds are formed from that part of alkene which is in between the double bonds.

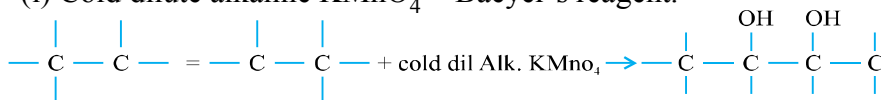


**For cyclic alkenes :**



(D) With potassium permagnate :

(i) Cold dilute alkaline  $\text{KMnO}_4$  = Baeyer's reagent.



[Alkene + cold dil.  $\text{KMnO}_4 \rightarrow$  Diol]

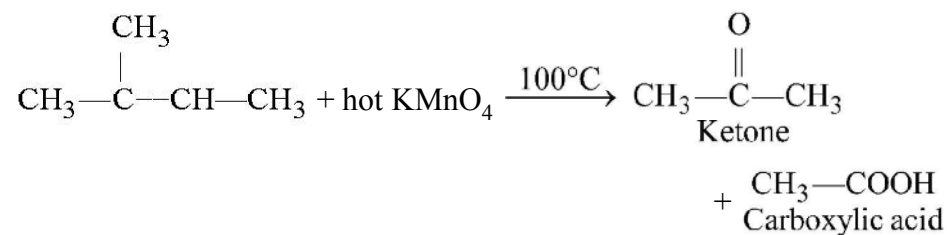
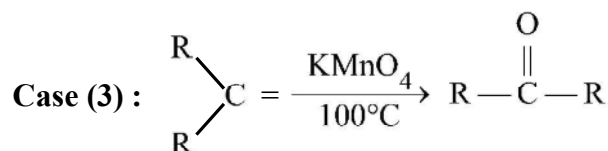
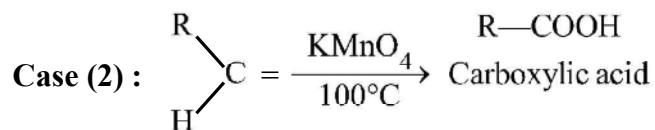
Baeyer's test for the presence of ( $\text{C}=\text{C}$ ) bond

Compound + cold dil. alk.  $\text{KMnO}_4 \longrightarrow$  Purple colour decolourised

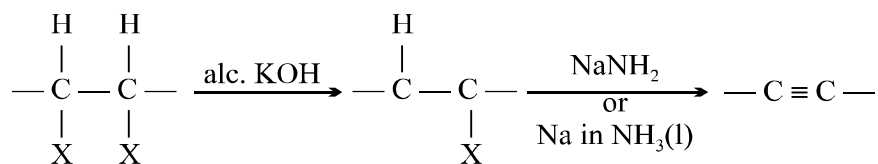
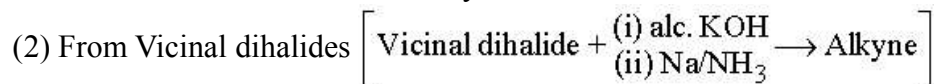
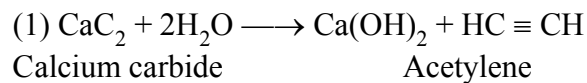
$\therefore$  Compound is alkene.



**Case (1) :**  $\begin{array}{c} \text{H} \\ \diagdown \\ \text{C} \\ \diagup \\ \text{H} \end{array} = \xrightarrow[100^{\circ}\text{C}]{\text{KMnO}_4} \text{CO}_2 + \text{H}_2\text{O}$

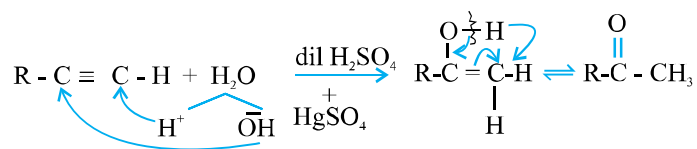
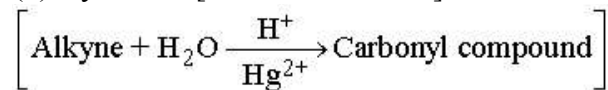


- **Preparation :**

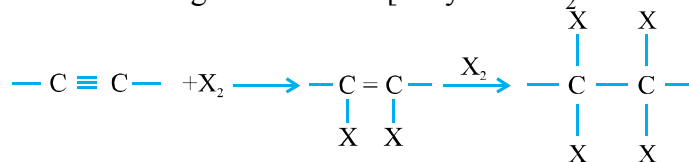


- **Chemical properties :**

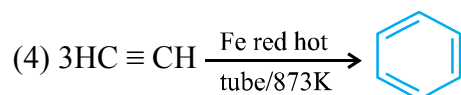
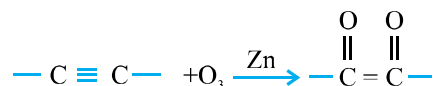
(1) Hydration [Addition of water]



(2) Addition of Halogen molecule : [Alkyne +  $2X_2 \longrightarrow$  Tetra halides]

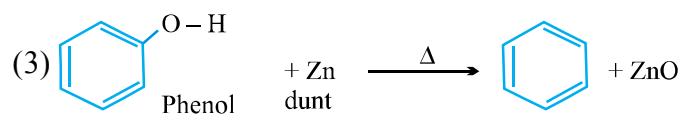
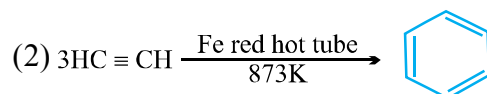
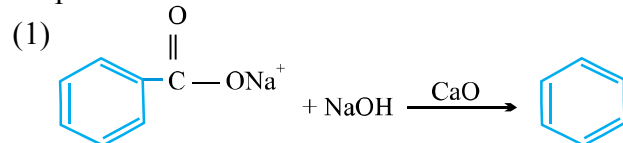


(3) Ozonolysis : [Alkyne +  $O_3 \longrightarrow$  Dicarbonyl compound]

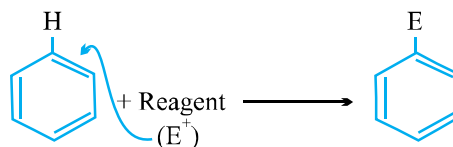


## Benzene

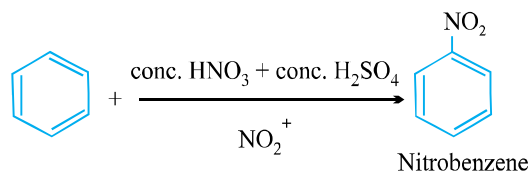
• Preparation :



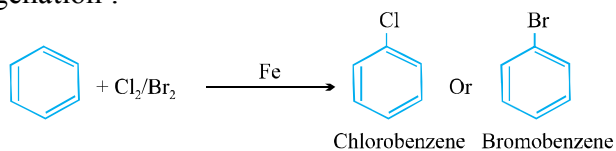
Chemical properties



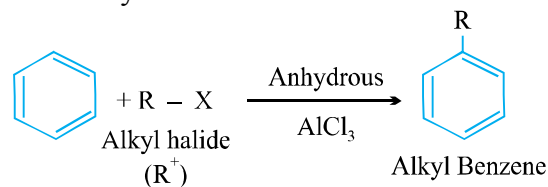
(1) Nitration :



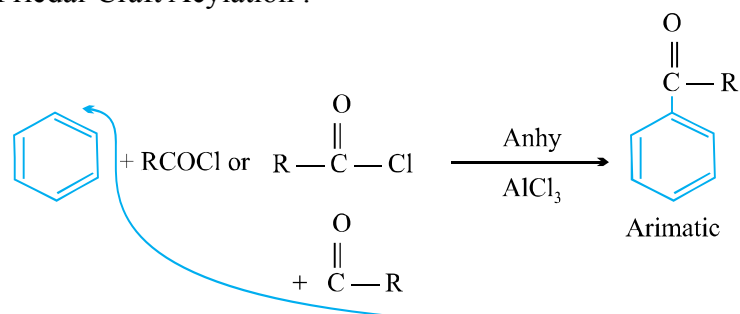
(2) Halogenation :



(3) Friedal-Craft Alkylation :



(4) Friedal-Craft Acylation :



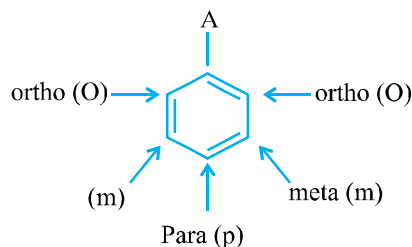
**Huckel's Rule :**

Conditions : (i) Compound must be planar.

(ii) Complete delocalisation of  $\pi e^-$

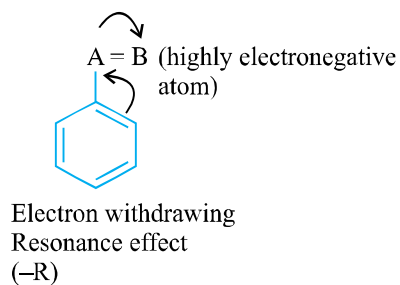
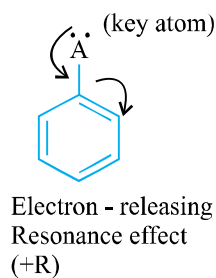
(iii) Presence of  $(4n + 2) \pi e^-$ . ( $n = 1, 2, 3, \dots$ )

**Derivatives of Benzenes :**

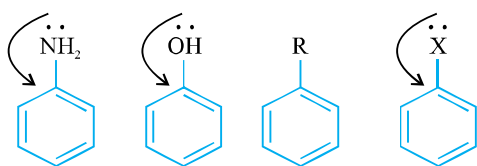


Ortho/Para directors : Group which direct the incoming electrophile to attach at ortho/para positions.

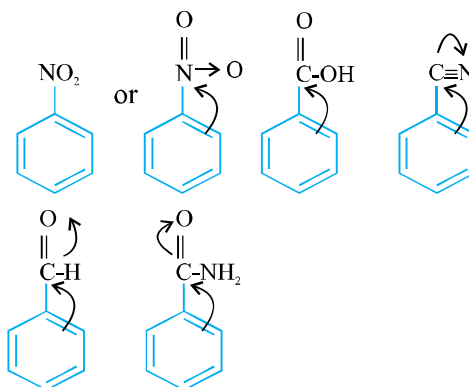
Meta directors : Groups which direct the incoming electrophile to attack at meta position.



Ortho/Para directions



Meta directions



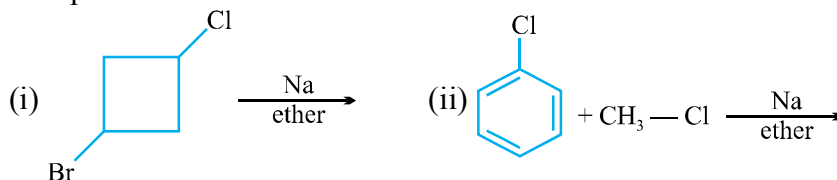
## Hydrocarbons

### 1-Mark Questions

- Give IUPAC name of  $\text{CH}(\text{CH}_3) - (\text{CH}_2)_4 - \text{CH}_3$
- Give the IUPAC name of
- Give the standard formula of 5-sec-butyl, 4-isopropyldecane.
- Give the standard formula of 4-tert-butyl, 4-ethyl-2, 2, 5, 5-tetra methyl hexane.
- Obtain isobutane from *n*-butane.
- $n\text{-Hexane} \xrightarrow[\text{V}_2\text{O}_5]{773\text{K}, 10-20 \text{ atm}} ?$

### 2-Mark Questions

- Out of 2-Methylpentane and 2, 3-Dimethylpentane which has greater boiling point and Why ?
- Give the structure of alkyl halide which when treated with sodium metal in presence of ether gives  $(\text{CH}_3)_2\text{CH}.\text{CH}(\text{CH}_3)_2$ .
- Complete :




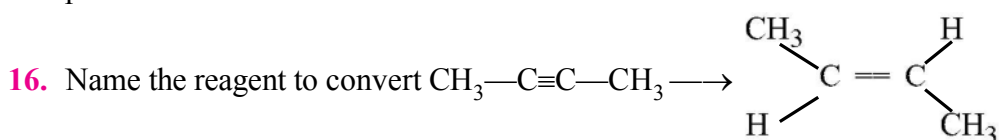
- Explain :

- Staggered form of ethane is more stable than eclipsed form.
- Wurtz reaction is carried out in dry ether.

## Alkenes

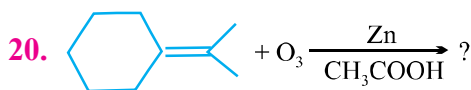
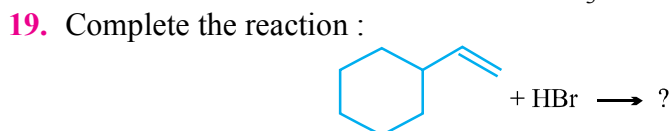
### 1-Mark Questions

11. Give IUPAC name of 
12. Give hybridisation of central carbon in allene ( $\text{CH}_2=\text{C}=\text{CH}_2$ )
13. Name the effect which decide the stability of alkenes.
14. Complete the reaction :  $\text{CH}_3-\text{CH}=\text{CH}_2 + \text{HCl} \xrightarrow{\text{Peroxide}} ?$
15. Which gas is produced during addition of HBr in alkenes in presence of peroxides ?

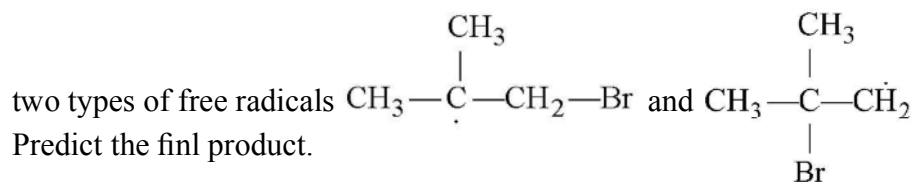


### 2-Mark Questions

17. Arrange the alkenes in decreasing order of stability.  
 $\text{CH}_3-\text{CH}=\text{CH}(\text{CH}_3)$ ,  $\text{CH}_2=\text{CH}_2$ ,  $\text{CH}_3-\text{CH}=\text{CH}_2$
18. Complete the reaction  $\text{CH}_2=\text{CH}-\underset{\text{CH}_3}{\text{C}}=\text{CH}_2 + \text{O}_3 \xrightarrow[\text{CH}_3\text{COOH}]{\text{Zn}} ?$



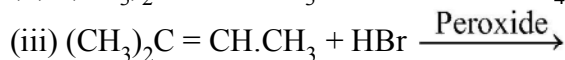
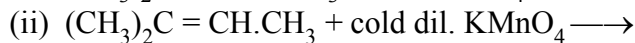
21. Name the alkene which will yield a mixture of cyclopentanone and propanal on treatment with  $\text{O}_3$  followed by reduction with Zn.
22. An alkene on treatment with  $\text{H}-\text{Br}$  in presence of peroxide can generate



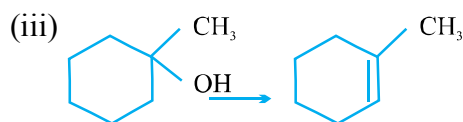
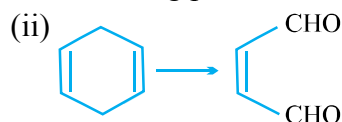
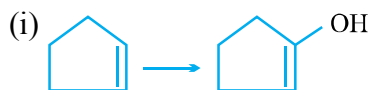
23. Explain :  
 (i) Melting point of *cis*-2-Butene is lower than that of *trans*-2-Butene.  
 (ii) Kharasch/peroxide effect is spontaneous with HBr only.

### 3 Mark Questions

24. Complete the reactions :



25. Indicate the reagents used to form the following products :



26. (i) Convert: iso-propylbromide  $\longrightarrow$  *n*-propyl bromide.

(ii) Give IUPAC name of Vinyl chloride.

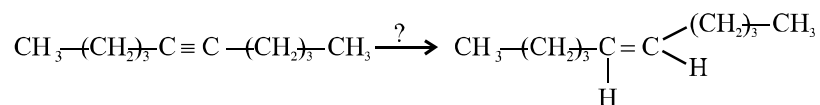
## Alkynes

### 1 Mark Questions

27. Give IUPAC name of acetylene.

28. Which alkyne would you start with to prepare  $\text{CH}_3-\text{CH}_2-\text{CH}_2-\text{CO}-\text{CH}_3$  ?

29. Name the reagent used in the following changes :

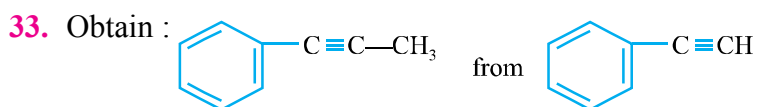


30. Give the alkyne which produce acetic acid and propanoic acid on treatment with alkaline  $\text{KMnO}_4$  at  $100^\circ\text{C}$ .

### 2 Mark Questions

31. Convert : Acetylene  $\rightarrow$  Propylene.

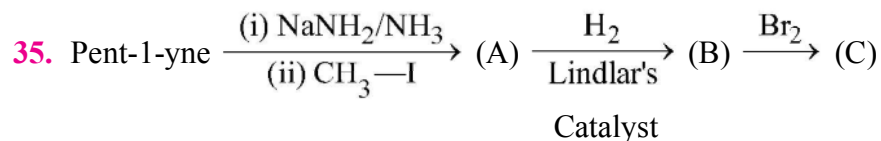
32. Convert : Ethylene  $\rightarrow$  Acetylene.



34. Give the product when 1-methylcyclohexane reacts with :

(i) aq. acidic  $\text{KMnO}_4$  (ii)  $\text{O}_3$  followed by  $\text{Zn}/\text{CH}_3\text{COOH}$ .

### 3 Mark Questions




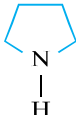
Identity A, B and C compounds and give their reactions.

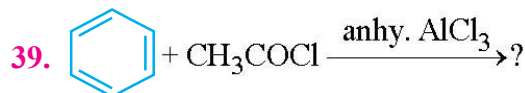
## Benzene

### 1 Mark Questions

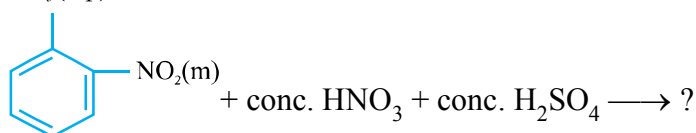
36. Who discovered benzene ?

37. Give reason whether  is aromatic or not.

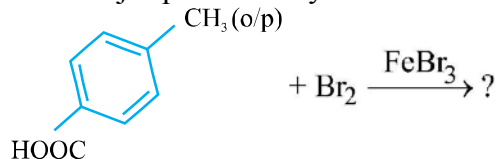
38. Is pyrrole  an aromatic compound or not ? Give reason.



40. Give major product only  
CH<sub>3</sub>(o/p)

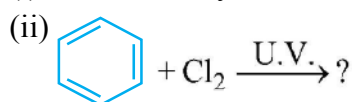


41. Give major product only



### 2 Mark Questions

42. (i) Convert Acetylene  $\longrightarrow$  Benzene



43. Distinguish chemically butyne and but-2-yne.

44. (i) Planar, cyclic, conjugated compounds with  $(4n + 2) \pi e^-$  are known as .....

(ii) Planar, cyclic, conjugated compounds with  $(4n) \pi e^-$  are known as .....

### 3 Mark Questions

45. Convert: Ethylene  $\longrightarrow$  Nitrobenzene.

46. Give chemical tests to distinguish the following :

- (i) Pent-1-yne and pent-2-yne
- (ii) Ethylene and Acetylene
- (iii) Ethane and Ethylene

47. Complete the following reactions :

- (i)  $\text{CH}_3\text{COONa} + \text{NaOH} \longrightarrow$
- (ii) *iso*-butyl bromide + alc. KOH  $\longrightarrow$
- (iii) *iso*-butyl alcohol + conc.  $\text{H}_2\text{SO}_4 \xrightarrow{\Delta}$
- (iv)  $(\text{CH}_3)_2\text{C}=\text{CH}.\text{CH}_3 + \text{HBr} \xrightarrow{\text{Peroxide}}$
- (v)  $\text{CH}_3-\underset{\text{CH}_3}{\text{C}}=\text{CH}_2 + \text{H}_2\text{O} \xrightarrow{\text{H}^+}$
- (vi)  $(\text{CH}_3)_2\text{C}=\text{CH}_2 + \text{cold dil. KMnO}_4 \longrightarrow$
- (vii)  $(\text{CH}_3)_2\text{C}=\text{CH}.\text{CH}_3 + \text{hot KMnO}_4 \longrightarrow$
- (viii)  $(\text{CH}_3)_2\text{C}=\text{CH}.\text{CH}_3 + \text{O}_3 \xrightarrow[\text{CH}_3\text{COOH}]{\text{Zn}}$
- (ix)  $\text{CH}_3-\underset{\text{Br}}{\text{CH}}-\underset{\text{Br}}{\text{CH}_3} + 2 \text{ alc. KOH} \longrightarrow$
- (x)  $\text{CaC}_2 + \text{H}_2\text{O} \longrightarrow$
- (xi)  $\text{CHCl}_3 + \text{Ag} \longrightarrow$
- (xii)  $\text{CH}_3-\text{C}\equiv\text{CH} + \text{H}_2\text{O} \xrightarrow[\text{Hg}^{2+}]{\text{H}^+}$
- (xiii)  $\text{CH}_3-\text{C}\equiv\text{CH}_2\text{CH}_3 + \text{O}_3 \xrightarrow[\text{H}_2\text{O}]{\text{Zn}}$
- (xiv)  $\text{CH}_3\text{C}\equiv\text{CH} + \text{alc. KMnO}_4 \xrightarrow{100^\circ\text{C}}$
- (xv)  $\text{C}_6\text{H}_6 + \text{H}_2 \xrightarrow[473 \text{ K}]{\text{Ni}}$

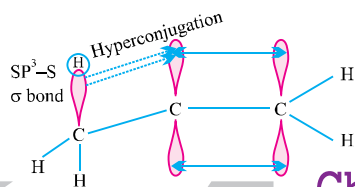
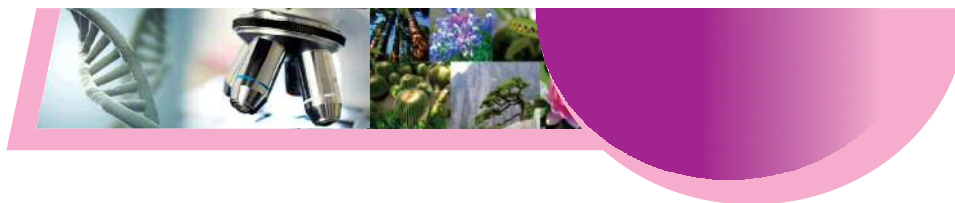
48. Conversions :

- (i) Ethane  $\rightarrow$  Ethyne
- (ii) Acetylene  $\rightarrow$  But-2-yne
- (iii) Propene  $\rightarrow$  Propanol



- (iv) Acetic acid  $\rightarrow$  Methane  
 (v) Acetylene  $\rightarrow$  Acetone ( $\text{CH}_3\text{COCH}_3$ )  
 (vi) Acetylene  $\rightarrow$  Cyclohexane

49. A hydrocarbon (X) on treatment with ammonical  $\text{AgNO}_3$  gave white precipitate. On treatment with water in dil.  $\text{H}_2\text{SO}_4$  and  $\text{HgSO}_4$ , it gave  $\text{CH}_3\text{—CHO}$ . When (X) is treated with 1 mol of  $\text{NaNH}_2/\text{NH}_3$ , along with *n*-propyl bromide, gave compound (Y), which on treatment with Lindlar's catalyst gave (Z) compound (Z) on treatment with  $\text{O}_3$  along with Zn gave  $\text{HCHO}$  and butanol. Identify X, Y, Z and give all the reactions.
50. An alkyl halide  $\text{C}_5\text{H}_{11}\text{Br}$  (A) reacts with alc. KOH to give an alkene (B) which reacts with  $\text{Br}_2$  to give compound (C), which on dehydrobromination gives an alkyne (D). On treatment with sodium metal in liquid ammonia, one mole of (D) gives one mole of sodium salt of (D) and half mole of  $\text{H}_2$  (g). Complete hydrogenation of (D) yields a straight chain alkane. Identify A, B, C and D. Give the reaction involved.
51. The sex attractant pheromone of codling moth has the molecular formula  $\text{C}_{13}\text{H}_{24}\text{O}$ . On catalytic reaction this compound gives 3-Ethyl-7-methyl-1-decanol having molecular formula  $\text{C}_{13}\text{H}_{28}\text{O}$ . On reduction ozonolysis the pheromone produces 2-pentane, 4-ketohexanal and 2-Hydroxyethanal. On the basis of these information, Write the structure of this pheromone.
52. 896 ml of a hydrocarbon (A) having 87.80% C and 12.19% H weights 3.28 g at STP. Hydrogenation of (A) gives 2-methylpentane. Also compound (A) on hydration in presence of  $\text{H}_2\text{SO}_4$  and  $\text{HgSO}_4$  gives Ketone (B) having molecular formula  $\text{C}_6\text{H}_{12}\text{O}$ . The ketone (B) gives a positive iodoform test. Find the structure of (A) give all reactions.
- [Hint : (i) 1 mole of a compound at STP contains 22400 mL volume  
 (ii) Ketones having  $\text{CH}_3\text{—}\overset{\text{O}}{\underset{\parallel}{\text{C}}}\text{—}$  structures gives positive iodoform test]
53. (a) Compound A  $\{\text{C}_{10}\text{H}_{18}\text{O}\}$  undergo reaction with  $\text{H}_2\text{SO}_4$  at  $250^\circ\text{C}$  to yield a mixture of two alkenes  $\{\text{C}_{10}\text{H}_{16}\}$ . The major alkene product (B) gives only cyclopentanone after ozone treatment followed by reduction with Zn in  $\text{CH}_3\text{COOH}$ . Identify (A) and (B) give the reactions involved.  
 (b) Convert  $\text{PhC}\equiv\text{CH} \longrightarrow \text{PhC}\equiv\text{C}\cdot\text{CH}_3$   
 (c) Benzene does not give addition reactions under normal conditions.



## Chapter - 14

# Environmental Chemistry

- **Environmental chemistry** deals with the study of the origin, transport, reactions, effects, fates of chemical species in the environment.
- **Environmental pollution** is the effect of undesirable changes in our surroundings that have harmful effects on plants, animals and human beings.
- **Main regions of atmosphere** are (i) troposphere (ii) Stratosphere (iii) mesosphere and (iv) ionosphere.
- **Pollutants** can be (i) gaseous air pollutants (*e.g.*,  $\text{SO}_2$ ,  $\text{NO}_2$ ,  $\text{CO}_2$ ,  $\text{H}_2\text{S}$ ,  $\text{O}_3$  hydrocarbons etc. and (ii) particular pollutants (*e.g.*, dust, mist, fumes, smoke, smog etc.)
- **Gaseous air pollutants** :  $\text{SO}_2$  can cause acute irritation to the membranes of the eyes resulting in tears and redness. It is also responsible for acid rain.  $\text{NO}_2$  is extremely toxic to living tissues, textiles and in the production of photochemical smog.
- **Particulate pollutants** : They are of two types :
  - (i) **Visible particulate or viable particulates** : These are the minute living organisms that are dispersed in atmosphere. Ex. Bacteria, fungi, moulds etc.
  - (ii) **Non-viable particulates** : These are formed either by the breakdown of larger materials or by the condensation of minute particles and droplets. For example mists, smoke, fumes and dust.
- **Green House Effect** is the phenomenon of warming of earth by absorption and re-emission of solar radiations.
- **Green house gases** are  $\text{CO}_2$ ,  $\text{CH}_4$ ,  $\text{O}_3$  water vapours, chlorofluoro carbons etc.
- **Acid rain** : When the pH of the rain water drops below 5.6, it is called acid rain. It damages monuments and structures made of firm marble, corrode articles made from metals, destroy plants and trees and it is also harmful to the aquatic life in lakes and rivers.
- **Ozone hole** is formed over South Pole due to depletion of ozone layer. CFCs and NO are responsible for ozone layer depletion.

- **Effects of depletion of ozone layer :** With the depletion of ozone layer, more UV radiation filters into troposphere. UV radiations lead to ageing of skin, cataract, sunburn, skin cancer, killing of many phytoplanktons, damage to fish productivity etc.
- **Smog** is derived from smoke and fog.
- **Classical smog** occurs in cool humid climate. It is a mixture of smoke, fog and sulphur dioxide. It is also called reducing smog.
- **Photochemical smog** occurs in warm and dry sunny climate. It has high concentration of oxidizing agents and therefore, it is also called as oxidizing smog.
- There are three main components of photochemical smog : nitrogen oxides, ozone and organic derivatives such as acrolein, formaldehyde, peroxyacetyl nitrate (PAN). PAN has the highest toxicity to plants attacking younger leaves and causing 'bronzing' and 'glazing' of their surface.
- **Water Pollution :** It is mainly caused by industrial waste which include heavy metals like Cd, Pb and Hg.
- **Eutrophication :** The process is which nutrient enriched water bodies support a dense plant population, which kills animal life by depriving it of oxygen and result in subsequent loss of biodiversity is known as Eutrophication.
- **BOD :** The amount of oxygen required by bacteria to break down the organic matter present in a certain volume of a sample of water, is called Biochemical Oxygen Demand (BOD).
- **Land Pollution :** It is caused by pesticides. Most pesticides can be divided into three categories—Insecticides, Herbicides and fungicides.
- The environmental pollution can be controlled :
  - (a) By recycling of household and industrial wastes.
  - (b) By sewage treatment.
  - (c) Incineration converts organic material to  $\text{CO}_2$  and  $\text{H}_2\text{O}$ .
- **Green Chemistry** is a way of thinking and is about utilizing the existing knowledge and principles of chemistry and other science to reduce the adverse impact on environment.
- Green Chemistry in Day to Day life :
  - (1) Dry cleaning of clothes by using liquified  $\text{CO}_2$  along with detergent instead of tetrachloroethene.
  - (2) Bleaching of paper and clothes by using  $\text{H}_2\text{O}_2$  a better bleaching agent as compared to  $\text{Cl}_2$ .

## Environmental chemistry

### 1-Mark Questions

1. Define environmental pollution.
2. Name three toxic metals that can pollute the environment.
3. Name the four non-viable particulates present in atmosphere.
4. Give one advantage and one disadvantage of ozone in atmosphere.
5. What is the name of the compound formed when CO combines with blood?
6. Which gas caused Bhopal Gas Tragedy? Give its formula.
7. What is the meaning of the term eutrophication with regard to water pollution?
8. Every year some people die by being in a room containing a faulty heater that uses coal, gas or oil. How might the death occur?
9. Mention two ways to reduce air pollution caused by automobiles.
10. How fluoride in tooth paste protects teeth against decay?  
[Hint : By converting hydroxyapatite (enamel on the surface of the teeth) into much harder fluorapatite]
11. What do you mean by Biochemical oxygen demand (BOD)?
12. Greenhouse effect leads to global warming. Which substances are responsible for greenhouse effect? What does CFC stand for?
13. What does CFC stand for?
14. Which out of  $\text{CO}_2$  and CO is more toxic and why?
15. Name the various components into which atmosphere can be divided.

### 2-Mark Questions

1. Write the adverse effect of excessive use of (i) fertilizers, and (ii) pesticides in the soil.
2. Write down the reactions involved during the formation of photochemical smog.
3. Why does rain water normally have a pH of about 5.6? When does it become acid rain? Why is acid rain considered as a threat to Taj Mahal?
4. What are biodegradable and non-biodegradable pollutants?
5. Explain giving reasons "The presence of CO reduces the amount of haemoglobin available in the blood for carrying oxygen to body cells."
6. What should be the tolerable limit of fluoride ions in drinking water? What happens if it is higher than 10 ppm?
7. Name four methods for waste management.
8. "Oxygen plays a key role in the troposphere while ozone in the stratosphere." Explain.

9. Write the effect of the following gases on human being :  $\text{SO}_2$ ,  $\text{NO}_2$ ,  $\text{CO}$ ,  $\text{CO}_2$ .
10. What are viable and non-viable particulates ?
11. How does  $\text{SO}_2$  cause pollution ? Give steps to control it.
12. A person started using underground water after facing acute shortage in municipality water supply. He felt taxative effect. What could be the cause ?

### 3-Mark Questions

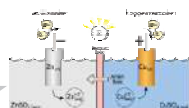
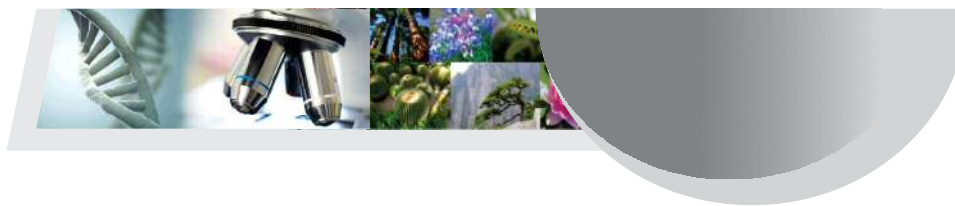
1. What do you understand by Ozone hole ? What are its consequences ?
2. What do you mean by green chemistry ? How will it help decrease environmental pollution ?
3. How can domestic waste be used as manure ?
4. What is the composition of photochemical smog and classical smog ? How do the two differ in their behaviour ?
5. A large no. of fish are suddenly found floating dead on a lake. There is no evidence of toxic dumping but you find an abundance of phytoplankton. Suggest a reason for the fish kill.
6. Do as directed :
  - (i) Name two important sinks of  $\text{CO}_2$ .
  - (ii) What is marine pollution ?
  - (iii) What is humification ?

[Hint : Production of humus by micro-organism in soil]
7. In view of green chemistry name the following :
  - (i) A chemical which can be used in place of CFC as blowing agent.
  - (ii) A chemical which can replace chlorine containing bleaching in paper industry.
8. Write the causes of depletion of ozone in the stratosphere. Write reactions also.
9. What are the harmful effects of small size particulate matters ?
10. How oxides of nitrogen are generated ? What are the harmful effects of oxides of nitrogen ?

### 5-Mark Questions

1. What do you understand by greenhouse effect ? What are the major greenhouse gases ? Why does green house effect leads to global warming ? What could be the consequences of global warming ?
2. Discuss the importance of dissolved oxygen in water. What processes are generally responsible for the deoxygenation of the water ?

3. How is photochemical smog formed ? What are harmful effects of photochemical smog ? Give some control measures to reduce photochemical smog ?
4. What is soil pollution ? Mention four ways of controlling soil pollution.
5. Explain the following terms :
  - (a) Green house effect
  - (b) Green chemistry
  - (c) Ozone layer depletion
  - (d) Global warming
  - (e) Eutrophication
6. (a) Discuss the effects if green house gases were totally missing in earth's atmosphere.  
(b) Chlorine radicals play an important role in ozone layer depletion. Write chemical reactions in support of the statement.



## Practicle Paper - I

# Chemisry

### Practice Paper-1

**Class : XI—Chemistry (Theory)**

**Time Allowed : 3Hrs**

**M.M. : 70**

#### General Instructions :

- All questions are compulsory.
- Q. No. 1 to 5 are very short answer type questions and carry one mark each.
- Q. No. 6 to 10 are very short answer type questions and carry two marks each.
- Q. No. 11 to 22 are short answer type questions and carry three marks each.
- Q. No. 23 is value based questions carries four marks.
- Q. No. 24 to 26 are long answer type questions and carry five marks each.
- Use log tables, if necessary, use of calculator is not allowed.

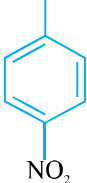
- Define limiting reagent 1
- Write the general electronic configuration of *d*-block elements ? 1
- Write vander Waal's equation for *n* moles of '*a*' real gas. 1
- What is the state of hybridisation of each carbon atom in C<sub>6</sub>H<sub>6</sub> ? 1
- Predict the sign of the entropy change ( $\Delta S$ ) for the following : 1  

$$\text{H}_2\text{O}(g) \longrightarrow \text{H}_2\text{O}(l)$$
- (i) Define Pauli's exclusion principle. 2  
 (ii) Write electronic configuration of an atom having 2K, 8L, 5M electrons. 2

**Or**

The mass of an electron is  $9.1 \times 10^{-31}$  kg. If its kinetic energy is  $3.0 \times 10^{-25}$  J. Calculate its wavelength. (Given  $h = 6.626 \times 10^{-34}$  Js)

- The concentration of hydrogen ion in a sample of soft dark is  $3.8 \times 10^{-3}$  M. What is its pH ? (log 3.8 = 0.58) 2

8. What happens when (write equation only) : 2  
 (i) Beryllium carbide reacts with water.  
 (ii) Sodium sulphate solution is added to an aqueous solution of barium nitrate.
9. Propanal and Pentan-3-one are the ozonolysis product of an alkene. What is the structural formula of the alkene ? 2
10. Write reactions to justify amphoteric nature of aluminium. 2
11. (a) Write the IUPAC name of the following : 3  
 (i)  (ii)  $\text{CH}_3 - \underset{\text{CH}=\text{CH}_2}{\text{CH}} - \text{CH}_2 - \text{CH}=\text{CH}_2$
- (b) Which isomerism is shown by following pair of compounds :  
 $\text{CH}_3\text{COOH}$  and  $\text{HCOOCH}_3$
12. Account for the following : 3  
 (a) An anion is always bigger than its parent atom.  
 (b) Chlorine (Cl) has more negative electron gain enthalpy than fluorine (F).  
 (c) Noble gases have positive electron gain enthalpy.
13. Give the shapes of following covalent molecules using VSEPR theory : 3  
 (i)  $\text{PCl}_5$  (ii)  $\text{BrF}_3$  (iii)  $\text{H}_2\text{O}$ .
14. A sample of drinking water was found to be severely contaminated with chloroform  $\text{CHCl}_3$ , supposed to be carcinogen. The level of contamination was 15 ppm (by mass). 3  
 (i) Express this in percent by mass.  
 (ii) Determine the mobility of chloroform in the water sample.
15. (a) State Hess's law. 3  
 (b) Use standard enthalpies of formation, calculate the value of  $\Delta_r H^\ominus$  for the reaction.  

$$2\text{H}_2\text{S}(\text{g}) + 3\text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) + 2\text{SO}_2(\text{g})$$

$$\Delta_f H^\ominus[\text{H}_2\text{O}(\text{l})] = -285.83 \text{ kJ/mol}$$

$$\Delta_f H^\ominus[\text{H}_2\text{S}(\text{g})] = -21.17 \text{ kJ/mol}$$



$$\Delta_f H^\ominus[\text{SO}_2(\text{g})] = -296.86 \text{ kJ/mol}$$

- 16.** (a) The molecular orbital theory to predict why the BE molecule does not exist.
- (b) Compare the stability of  $\text{O}_2^+$  and  $\text{O}_2$  on the basis of M.O. theory. **3**
- 17.** (a) Calculate the total pressure in a mixture of 8g of oxygen and 4g of hydrogen confined in a vessel of  $1 \text{ dm}^3$  at  $27^\circ\text{C}$ .
- [Given  $R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$ ] **3**
- (b) Critical temperature of  $\text{CO}_2$  and  $\text{CH}_4$  are  $31.1^\circ\text{C}$  and  $-81.9^\circ\text{C}$  respectively. Which of there has stronger intermolecular forces and why ?

***Or***

- (a) What would be the S.I. units of a quantity  $PV^2T^2/n$  ?
- (b) Calculate the temperature of 4.0 moles of a gas occupying  $5 \text{ dm}^3$  at 3.32 bar ( $R = 0.083 \text{ bar dm}^3\text{K}^{-1} \text{ mol}^{-1}$ )
- 18.** (a) What is the oxidation no. of Cr. in  $\text{Cr}_2\text{O}_7^{2-}$  ? **3**
- (b) Balance the equation :
- $$\text{MnO}_4^- + \text{Fe}^{2+} \longrightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}^+ \quad (\text{Acidic medium})$$
- 19.** Complete the following chemical reactions. **3**
- (a)  $\text{PbS}(s) + \text{H}_2\text{O}_2(aq) \longrightarrow$
- (b)  $\text{MnO}_4^-(aq) + \text{H}_2\text{O}_2(aq) \longrightarrow$
- (c)  $\text{Ca}_3\text{N}_2 + \text{H}_2\text{O} \longrightarrow$
- 20.** Give reason for the following : **3**
- (a)  $\text{PbCl}_4$  is a powerful oxidising agent.
- (b) Graphite acts as a good lubricant.
- (c) Boron halides do not dimerise like  $\text{BH}_3$
- 21.** (a) Draw the structure of (i)  $\text{BeCl}_2$  (vapour) and (ii)  $\text{BeCl}_2$  solid
- (b) Complete the equation :
- $$\text{Cl}_2 + \text{Ca}(\text{OH})_2 \longrightarrow$$

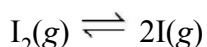
22. Arrange the following in order of property mentioned against each : 3
- (a)  $(\text{CH}_3)_3\text{C}^\oplus$ ,  $\text{CH}_3 - \text{CH}_2 - \underset{\text{CH}_3}{\overset{\oplus}{\text{CH}}}$ ,  $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2^\oplus$   
(increasing order of stability)
- (b)  $\text{HCOOH}$ ,  $\text{CH}_3\text{COOH}$ ,  $\text{CH}_3\text{CH}_2\text{COOH}$  (increasing acidic strength)
- (c)  $\dot{\text{C}}\text{H}_3$ ,  $\text{CH}_3 - \dot{\text{C}}\text{H} - \text{CH}_3$ ,  $\text{CH}_3 - \underset{\text{CH}_3}{\overset{\text{CH}_3}{\text{C}}}$  (increasing order of stability)

### Valued Based Questions

23. Environmental pollution is causing a serious threat on the earth. Due to combustion of fossil fuels, a number of poisonous and harmful gases enter into the atmosphere. Every country has made strict laws for industries and individual citizens to keep pollution under control. 4

Now answer the following questions :

- (a) Why is it advised not to sleep with burning coke anywhere in a closed room on winter nights ?
  - (b) What does the combustion of motor fuels cause pollution of the atmosphere ?
  - (c) Name one natural source and one human activity by which  $\text{SO}_2$  enters into atmosphere ?
24. (a) Explain the following with example : 5
- (i) Common ion effect
  - (ii) Buffer solution.
- (b) At a certain temperature and total pressure of  $10^5$  Pa, iodine vapours contain 40% by volume of I atoms.



Calculate  $K_p$  for the equilibrium.

**Or**

- (a) Define Le-Chatelier's principle.
  - (b) Define pH.
  - (c) Equilibrium constant for the reaction is 4.0. What will be the equilibrium constant for the reverse reaction.
  - (d) Calculate the pH of  $10^{-8}$  M HCl solution.
25. (a) Define Heisenberg's uncertainty principle. Write its mathematical expression. 5
- (b) Calculate the uncertainty in the velocity of a cricket ball of mass 150 g. If the uncertainty in its position is of the order of  $1 \text{ \AA}$ .  
( $h = 6.6 \times 10^{-34} \text{ Kg m}^2\text{s}^{-1}$ )

**Or**

- (a) Which of the following orbitals are not possible ?  
 $1p, 2s, 2p, 3s$
- (b) Which of the following sets of quantum number are not possible ? Give reasons.
- (i)  $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$
  - (ii)  $n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$

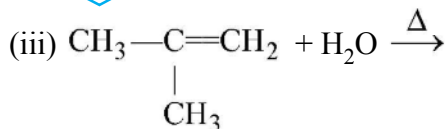
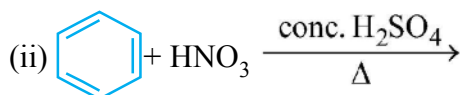
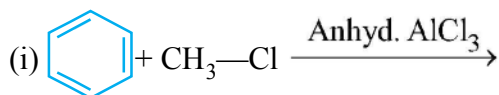
- (c) Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength  $6800\text{\AA}$ . Calculate threshold frequency ( $\nu_0$ ) and work function ( $\phi_0$ ) of the metal.

26. (a) Define the following with example :

5

- (i) Wurtz reaction  
(ii) Markovnikov's Rule.

(b) Give the main product of the reaction :

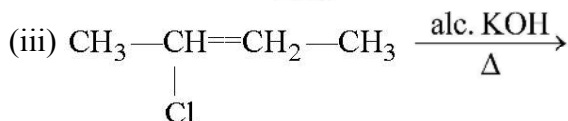
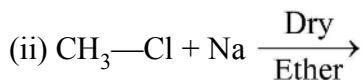
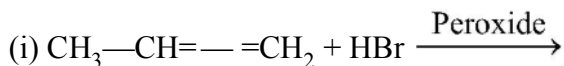


*Or*

(a) How can you convert the following :

- (i) Propan-1-ol to propan-2-ol  
(ii) Ethyl chloride to *n*-Butane

(b) Complete the following reactions :



## Practice Paper-1

### Solution

1. The reactant which is completely consumed during the reaction is called limiting reagent. 1
2.  $(n-1)d^{1-10} ns^{1-2}$ . 1
3.  $\left[ P + \frac{an^2}{V^2} \right] (V - nb) = nRT$  1
4.  $sp^2$  hybridisation. 1
5. – ve, as randomners decrease. 1
6. (a) No two  $e^-$  in an atom can have same set of all the four quantum numbers. 1 + 1

(b) Total  $e^- = 15$

$$\text{E.C.} = 1s^2 2s^2 2p^6 3s^2 3p^3$$

**Or**

$$\text{K.E.} = \frac{1}{2}mv^2$$

$$V = \left( \frac{2 \times \text{K.E.}}{m} \right)^{1/2} = \left( \frac{2 \times 3 \times 10^{-25} \text{ kg m}^2 \text{ s}^{-2}}{9.1 \times 10^{-31} \text{ kg}} \right)^{1/2}$$

$$V = 8.12 \times 10^2 \text{ m/s} = 812 \text{ m/s} \quad 1$$

$$\lambda = \frac{h}{mv}$$

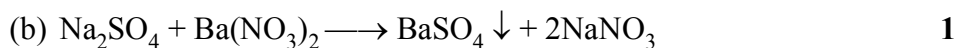
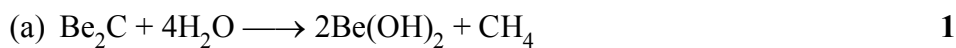
$$= \frac{6.626 \times 10^{-34} \text{ Js}}{(9.1 \times 10^{-31} \text{ kg})(812 \text{ ms}^{-1})} = 896.7 \text{ nm} \quad 1$$

$$\text{pH} = -\log [\text{H}^+] \quad \frac{1}{2}$$

$$= -\log(3.8 \times 10^{-3}) \quad \frac{1}{2}$$

$$= -\log 3.8 = 3 - 0.58$$

$$= 2.42 \quad 1$$



9.  $\text{CH}_3-\text{CH}_2-\text{CH}=\text{O} + \text{O}=\text{C}\begin{matrix} \text{CH}_2-\text{CH}_3 \\ \text{CH}_2-\text{CH}_3 \end{matrix}$   $\frac{1}{2} + \frac{1}{2}$
- $\text{CH}_3-\text{CH}_2-\text{CH}=\text{C}\begin{matrix} \text{CH}_2-\text{CH}_3 \\ \text{CH}_2-\text{CH}_3 \end{matrix}$  1
10.  $\Delta H^\circ = \Delta E^\circ + \Delta n g R T$
- $\Delta H^\circ = -10500 \text{ J} - 2477.57 \text{ J}$
- $\Delta H^\circ = -12977.57 \text{ J/mol}$  1
- $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$
- $= -12977.57 + 13141.8$   $\frac{1}{2}$
- $\Delta G^\circ = 0.164 \text{ kJ}$ , Process is non-spontaneous  $\frac{1}{2}$
11. (a) (i) 4-Nitro benzoic acid 1
- (ii) 3-Methyl hexa-1, 5-diene 1
- (b) Functional isomerism 1
12. (a) Because nuclear charge decrease/ $e^-$  increases. 1
- (b) Because of bigger size of chlorine/less  $e^- - e^-$  repulsion as compared to fluorine. 1
- (c) Because they have fully filled orbitals. 1
13. (a) Trigonal bipyramidal. 1
- (b) Bent T-shape. 1
- (c) Bent/V-shape. 1
14. (a) 15 ppm means 15 parts in million ( $10^6$ ) parts. a
- $\therefore \% \text{ by mass} = \frac{15}{10^6} \times 100 = 15 \times 10^{-4} = 1.5 \times 10^{-3} \%$  1
- (b) Molar mass of  $\text{CHCl}_3 = 119.5 \text{ g/mol}$
- 100 g of sample contain chloroform  $= 1.5 \times 10^{-3} \text{ g}$
- $\therefore 1000 \text{ g of sample contain} = \frac{1.5 \times 10^{-3}}{100} \times 100 = 1.5 \times 10^{-2} \text{ g}$  1
- Molality of chloroform  $= \frac{1.5 \times 10^{-2}}{119.5 \times 1} = 1.255 \times 10^{-4}$  1
- $\therefore \text{Molality} = 1.255 \times 10^{-4} \text{ m}$
15. (a)  $2\text{Al}(s) + 3\text{H}_2\text{SO}_4(aq) \longrightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 3\text{H}_2(g)$  1
- (b)  $2\text{Al}(s) + 2\text{NaOH}(aq) + 6\text{H}_2\text{O}(l) \rightarrow 2\text{Na}^+[\text{Al}(\text{OH})_4]^- (aq) + 3\text{H}_2(g)$
- Sodium tetrahydroxo aluminate (III) 1

16. (a) Be :  $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2$

$$\text{Bond order} = \frac{N_b - N_a}{2} = \frac{4 - 4}{2} = 0 \quad 1$$

As bond order of  $\text{Be}_2$  is zero and does not exist.

(b)  $\text{O}_2$  : B.O =  $\frac{N_b - N_a}{2}$  1/2

$$= \frac{10 - 6}{2} = 2$$

$\text{O}_2^+$  : B.O =  $\frac{N_b - N_a}{2}$  1/2

$$= \frac{10 - 5}{2} = 2.5$$

$$\text{O}_2^+ > \text{O}_2$$

Greater the bond order, more will be stability. 1

17. (a)  $n_{\text{O}_2} = \frac{8}{32} = 0.25 \text{ mol}$

$$n_{\text{H}_2} = \frac{4}{2} = 2 \text{ mol}$$

Total no. of moles =  $2 + 0.25 = 2.25 \text{ mol}$  1

$$V = 1 \text{ dm}^3, T = 300 \text{ K}, R = 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$$

$$PV = nRT$$

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

$$= \frac{2.25 \text{ mol} \times 0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{1 \text{ dm}^3} \quad 1$$

$$P = 56.025 \text{ bar}$$

(b)  $\text{CO}_2$  has stronger intermolecular forces because it can be liquefied at temperature upto  $31.1^\circ\text{C}$  but  $\text{CH}_4$  can be liquefied only upto  $-81.9^\circ\text{C}$ .

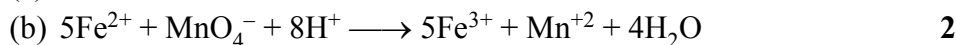
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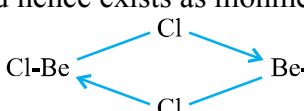
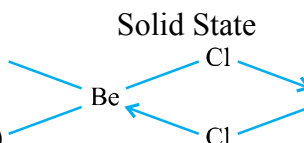
$$\frac{PV^2T^2}{V} = \frac{(\text{Nm}^{-2})(\text{m}^3)^2(\text{K}^2)}{\text{mol}} = \text{Nm}^4 \text{ K}^2 \text{ mol}^{-1} \quad 1$$

$$PV = nRT \text{ or } T = \frac{PV}{nR} \quad 1$$

$$T = \frac{3.32 \times 5}{4 \times 0.083} = 50 \text{ K} \quad 1$$

18. (a)  $x = 6$



19. (a)  $\text{PbS} + 4\text{H}_2\text{O}_2 \longrightarrow \text{PbSO}_4 + 4\text{H}_2\text{O}$  1  
 (b)  $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 \longrightarrow 2\text{Mn}^{+2} + 8\text{H}_2\text{O} + 5\text{O}_2$  1  
 (c)  $\text{Ca}_3\text{N}_2 + 6\text{H}_2\text{O} \longrightarrow \text{Ca}(\text{OH})_2 + 2\text{NH}_3$  1
20. (a) Because  $\text{Pb}^{+4}$  is reduced to more stable  $\text{Pb}^{+2}$ . 1  
 (b) Because of its soft and slippery nature. 1  
 (c) Boron due its small size cannot four large size chlorine atoms around it and hence exists as monmer. 1
21. (a) (i)  1  
 (ii)  1
- (b)  $\text{Cl}_2 + 2\text{Ca}(\text{OH})_2 \longrightarrow \text{Ca}(\text{OCl})_2 + \text{H}_2\text{O} + \text{CaCl}_2$  1  
 Slaked lime Bleaching  
 powder
22. (a)  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2^+ < \text{CH}_3\text{CH}_2\text{CH}^+ < (\text{CH}_3)_3\text{C}^+$  1  
 CH<sub>3</sub>  
 Increasing acidic strength
- (b)  $\text{CH}_3\text{CH}_2\text{COOH} < \text{CH}_3\text{COOH} < \text{HCOOH}$  1  
 increasing acidic strength
- (c)  $\text{CH}_3 < \text{CH}_3\text{CHCH}_3 < \text{CH}_3\text{C}(\text{CH}_3)_2$  1  
 increasing order of stability
23. (a) On burning coke, a lot of CO is produced which causes anoxia. 1  
 (b) On burnign motal-fuels, toxic oxides of nitrogen is formed which affect respiratory system. 1  
 (c) Due to volcanic eruption and through combustion of sulphur containing fuels. 1  
 (d) Concern for environment. 1

24. (a) (i) It is the shift in an ionic equilibrium caused by the addition of a solute that provides an ion that takes part in the equilibrium.

(ii) The solution which resists the change in pH on dilution/addition of acid or alkali. 1

(b) Partial pressure of I atom ( $P_I$ ) =  $\frac{40}{100} \times 10^5 = 0.4 \times 10^5$  Pa 1

Partial pressure of  $I_2$  atom ( $P_{I_2}$ ) =  $\frac{60}{100} \times 10^5 = 0.6 \times 10^5$  Pa 1

$$K_p = \frac{[P_I]^2}{[P_{I_2}]} = \frac{(0.4 \times 10^5)^2}{(0.6 \times 10^5)} = 2.67 \times 10^4 \text{ Pa} \quad 1$$

(i) Correct statement. 1

(ii) negative logarithm of hydrogen ion concentration. 1

(iii)  $\frac{1}{4}$  1

(iv)  $10^{-8}$  M HCl

$$[H^+] = 10^{-8} + 10^{-7} \quad 2$$

[ $\therefore$   $[H^+]$  conc. due to water =  $10^{-7}$ ]

or  $[H^+] = 11 \times 10^{-8} \Rightarrow PH = -\log[11 \times 10^{-8}]$   
 $= 8 - 1.02 = 6.98$

25. (a) It is impossible to measure simultaneously both the position and velocity of a microscopic particle with accuracy or certainty. 1 + 1

Mathematically expression  $\Delta x \times \Delta p \geq \frac{h}{4\pi}$

(b) Mass of ball  $m = 150$  g

$$= 150 \times 10^{-3} \text{ kg} = 0.150 \text{ kg}$$

Uncertainty in position  $\Delta x = 1 \text{ \AA} = 10^{-10} \text{ m}$

$$\Delta x \cdot \Delta V \times m = \frac{h}{4\pi} \quad 1$$

$$\Delta V = \frac{h}{4\pi \times \Delta x \times m}$$

$$= \frac{6.626 \times 10^{-34} \text{ kg m}^2 \text{ s}^{-1}}{4 \times 3.14 \times 10^{-10} \text{ m} \times 0.150 \text{ kg}} \quad 1$$

$$= 3.52 \times 10^{-24} \text{ m s}^{-1} \quad 1$$



**Or**

(a)  $1p$  as  $n = l$ , not possible.

(b) (i)  $n = 0, l = 0, m_l = 0, m_s = +\frac{1}{2}$  not possible because  $n$  cannot be zero. 1

(c) Energy of photon  $= \frac{hc}{\lambda}$   
 $= \frac{6.63 \times 10^{-34} \text{ Js} \times 3 \times 10^8 \text{ ms}^{-1}}{6800 \times 10^{-10} \text{ m}}$  1

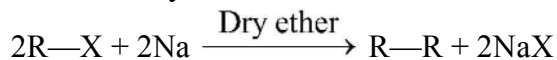
$= 2.93 \times 10^{-19} \text{ J}$

Energy of photon = K.E. +  $W_0 = 0 + W_0 = W_0$  1

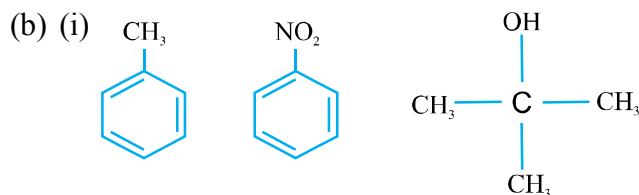
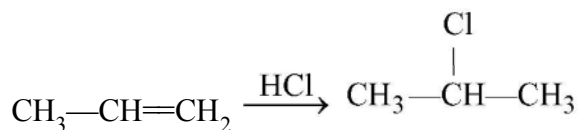
$W_0 = 2.93 \times 10^{-19} \text{ J}$

Threshold frequency  $\nu_0 = \frac{c}{\lambda_0} = \frac{3 \times 10^8 \text{ ms}^{-1}}{6800 \times 10^{-10} \text{ m}} = 4.41 \times 10^{14} \text{ s}^{-1}$  1

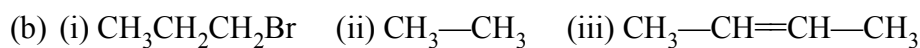
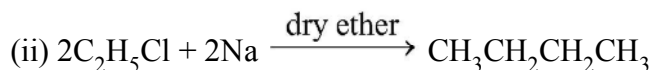
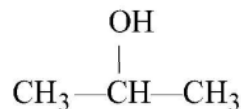
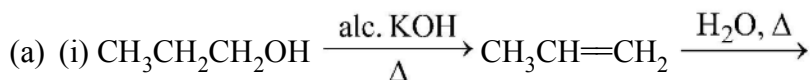
26. (a) (i) It involves the chemical reaction between alkyl halides and metallic sodium in presence of dry ether and from alkanes. 1



(ii) It states that 'more' electronegative part of the addendum adds to that carbon of double bond which contains lesser no. of H-atoms.



**Or**



**Practice Paper-1**  
**Class : XI**  
**Blue Print**

S.No.	Name of Chapter	V.S.A.	S.A.	S.A.	L.A.	Total Marks
1.	Some Basic Concepts of Chemistry	1		1		04
2.	Structure of Atom		1		1	07
3.	Classification of Elements	1		1		04
4.	Chemical Bonding			2		06
5.	States of Matter	1		1		04
6.	Thermodynamics	1		1		04
7.	Equilibrium		1		1	07
8.	Redox Reaction			1		03
9.	Hydrogen			1		03
10.	s-Block		1	1		05
11.	p-Block		1	1		05
12.	Organic Chemistry	1		2		07
13.	Hydrocarbon		1		1	07
14.	Environmental Chemistry				1 VBQ	04
					4 marks	
	<b>Total</b>	05	05	12	3+1 (VBQ)	70

## Sample Question Paper

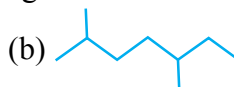
- (a) All questions are compulsory.
- (b) Q. No. 1 to 5 are Very Short Question carry one mark each.
- (c) Q. No. 6 to 10 are Short Answer Questions and carry 2 marks each.
- (d) Q. No. 11 to 22 are Short Answer Questions and carry 3 marks each.
- (e) Q. No. 23 is Value Based Questions carries 4 marks.
- (f) Q. No. 24 to 26 are Long Answer Questions and carry 5 marks each.

1. Which of the following has maximum number of significant figure ?  
(i) 0.00453 (ii) 4.8046 (iii) 5.643.
2. How are 0.5 m NaOH and 0.5 M NaOH different from each other ?
3. Explain why  $\text{Na}^+$  is smaller in size than Na atom.
4. Draw resonating structures of  $\text{CO}_3^{2-}$  ion.
5. Define critical temperature.
6. Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass. [Atomic mass : Fe = 55.85, O = 16.00]
7. (a) Write the electronic configuration of  $\text{Cu}^{2+}$  ion.  
(b) Why are Bohr's orbits called stationary states.
8. (a) How many subshells are associated with  $1H = 4$  ?  
(b) How many electrons will be present in the subshells having  $m_s$  value of  $-\frac{1}{2}$  for  $n = 4$  ?
9. (a) What do you understand by isoelectronic species ?  
(b) Write a cation and an anion which is isoelectronic with Ar.
10. Give the shapes of following covalent molecules using VSEPR theory :  
(i)  $\text{NH}_3$  (ii)  $\text{ClF}_5$  2

**Or**

- (a) Use molecular orbital theory to predict why  $\text{Be}_2$  molecule does not exist. 1 + 1
- (b) Compare the stability of  $\text{O}_2^+$  and  $\text{O}_2^-$ .
11. The work function for caesium atom is 1.9 eV. Calculate.  
(a) The threshold wavelength.  
(b) If the caesium element is irradiated with a wavelength of 500 nm, calculate the kinetic energy of ejected electron. [Given  $1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}$ ]  
1½ + 1½

12. Use the periodic table to identify the following :  
 (a) A group whose elements show valence 2 and 6. 1  
 (b) A metal which can form a predominantly stable covalent halide of the formula MX. 1  
 (c) The group having metalloid, non-metal liquid as well as gas at room temp. 1
13. Compare the relative stabilities of  $\text{O}_2^-$  and  $\text{N}_2^+$  and comment on their magnetic behaviour. 3
14. Give reasons for the following : 3  
 (i)  $\text{H}_2\text{O}$  has higher boiling point than HF.  
 (ii) Ice floats on the surface of water.  
 (iii)  $\text{KHF}_2$  exists but  $\text{KCl}_2$  does not.
15. What is the hybridisation of the central atom in (i)  $\text{H}_3\text{O}^+$ , (ii)  $\text{XeF}_2$ , (iii)  $\text{XeF}_4$ ? What are their shapes ? 3
16. (a) In terms of Charles's law, explain why  $-273^\circ\text{C}$  is the lowest temperature? 3  
 (b) 20 ml of hydrogen measured at  $15^\circ\text{C}$  are heated to  $35^\circ\text{C}$ , what is the new volume at the same pressure.
17. Define van der Waals' equation. The van der Waals' constants for two gases are as follows : 3
- | Gas | a (atm $\text{L}^2 \text{mol}^{-1}$ ) | b ( $\text{L mol}^{-1}$ ) |
|-----|---------------------------------------|---------------------------|
| X   | 1.39                                  | 0.0391                    |
| Y   | 3.59                                  | 0.427                     |
- Which of them more easily liquefiable and which has greater molecular size?
18. (a) Name the different salts that cause permanent hardness of water. (any two) 3  
 (b) How do we obtain demineralised water from hard water after passing it from synthetic ion exchange resins? Give reaction ?
19. What happens when : 3  
 (a) Boric acid is added to water.  
 (b) Al is treated with dil. NaOH.  
 (c)  $\text{BF}_3$  is treated with  $\text{NH}_3$
20. Write IUPAC names of the following : 3

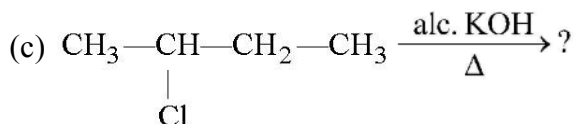
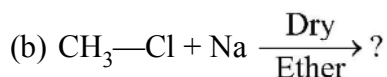
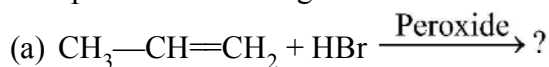


21. Write the naming reactions :

- (i) Wurtz reaction
- (ii) Decarboxylation of sodium salt of fatty acid
- (iii) Friedal Craft reaction.

*Or*

Complete the following reaction :



22. (a) Mention two similarities in the behaviour of Be and Al to show that they have diagonal relationship.

(b) What is the biological importance of Na in our body ?

23. A factory was started near a village. Suddenly villagers started feeling the presence of irritating vapours in the village and cases of headache, chestpain, cough and breathing problem increased. Anil a science student blamed the emission from the chimney of the factory for such problems ?

Reading the above paragraph, answer the following questions :

(a) Explain what could happen ?

(b) Give chemical reactions for the support of your explanation.

(c) Mention the values shown by student in paragraph.

24. (a) What happens when :

(i) Propene is treated with HBr.

(ii) Benzene is treated with methyl chloride in the presence of anhyd.  $\text{AlCl}_3$

(iii) When methyl bromide treated with sodium metal in presence of dry ether.

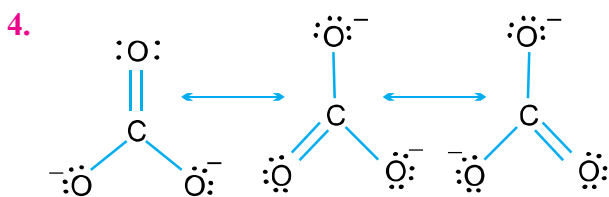
(b) Prove that phenolic group is ortho and para directing.

*Or*

- (a) How is benzene obtained from ethyne ? What is process called ?
  - (b) What happens when ethyne is heated with water in the presence of mercuric sulphate and dilute hydrochloric acid ?
  - (c) What is Markownikow's Kov rule ?
- 25.** (a) Equilibrium constant for a reaction is 10. What will be the equilibrium constant for the reverse reaction ?
- (b) Write the conjugate acids for the Bronsted base  $\text{OH}^-$  and  $\text{CH}_3\text{COO}^-$ .
  - (c) Determine the pH of  $10^{-8}$  M HCl solution taking into account the  $\text{H}^+$  produced by water also. ( $\log 11 = 1.0414$ ).
- 26.** (a) Propanal and pentan-3-one are the ozonolysis product of an alkene. What is the structural formula of the alkene ?
- (b) An Alkyl halide (A) of formula  $\text{C}_6\text{H}_{13}\text{Cl}$  on treatment with alcoholic KOH give two isomeric alkenes (B) and (C) ( $\text{C}_6\text{H}_{12}$ ). Both alkenes on hydrogenation give 2, 3-dimethyl butane. Predict the structure of A, B and C.

## Marking Scheme

1. 4.8046. 1
2. In 0.5 m NaOH, 1 mol NaOH present in 1 kg solvent while in 1.0 M NaOH, 1 mol NaOH present in 1 L of solution. 1
3.  $\text{Na}^+$  has only 2 shells and more effective nuclear charge. 1



5. Refer NCERT  
Appearance of gas at that temperature is called critical temperature  $\text{CO}_2 = -30.98^\circ\text{C}$ .

Element	% by mass	At mass	Relative	Simple	Whole no.
Fe	69.9	55.85	$\frac{69.9}{55.85} = 1.25$	1	2
O	30.1	16.00	$\frac{30.1}{16.00} = 1.88$	1.5	3

Empirical formula :  $\text{Fe}_2\text{O}_3$

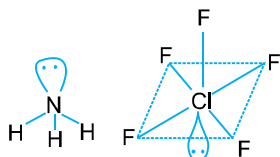
7. (a)  $\text{Cu}^{2+}$  (27)  $[\text{Ar}]3d^94s^0$ .  
(b) When an electron is revolving in Bohr's orbit it does not lose energy.
8. (a)  $n = 4, l = 0, 1, 2, 3$   
 $s, p, d, f$  subshells are associated with  $n = 4$ .  
(b) Total number of electrons in shell  $n = 4$  is 32 half of this *i.e.*, 16 elements will have  $m_s = -\frac{1}{2}$ .
9. (a) The species having the same number of electrons are called iso electric species.  
(b)  $\text{Ca}^{2+}, \text{Cl}^-$
10. (a) M.O. configuration of  $\text{Be}_2$  molecule is :  
 $\sigma 1s^2 \sigma^* 1s^2 \sigma 2s^2 \sigma^* 2s^2$   
Bond order  $= \frac{1}{2} (4 - 4) = 0$   
Since bond order is zero  $\text{Be}_2$  molecule is not possible.

$$(b) \text{B.O. (O}_2^+) = \frac{1}{2}(10 - 5) = 2.5$$

$$\text{B.O. (O}_2^-) = \frac{1}{2}(10 - 7) = 1.5$$

Since B.O. of  $\text{O}_2^+$  is more than that of  $\text{O}_2^-$ ,  $\text{O}_2^+$  is more stable.

(c)



$$11. (a) W_0 = \frac{hc}{\lambda_0}, \lambda_0 = \frac{hc}{W_0}$$

$$\lambda_0 = \frac{6.626 \times 10^{-34} \text{ Js} \times 3 \times 10^8 \text{ ms}^{-1}}{1.9 \times 1.6 \times 10^{-19} \text{ J}} = 6.538 \times 10^{-7} \text{ m}$$

$$(b) E = W_0 + \text{K.E.}$$

$$\frac{hc}{\lambda} = \frac{hc}{\lambda_0} + \text{K.E.}$$

$$\text{K.E.} = hc \left[ \frac{1}{\lambda} - \frac{1}{\lambda_0} \right]$$

$$= 6.626 \times 10^{-34} \times 3 \times 10^8 \left( \frac{1}{5 \times 10^{-7}} - \frac{1}{6.54 \times 10^{-7}} \right)$$

$$= 19.878 \times 10^{-19} \left( \frac{1}{5} - \frac{1}{6.54} \right)$$

$$= 19.88 \times 10^{-19} (0.2 - 0.15)$$

$$= 0.994 \times 10^{-19} \text{ J}$$

12. (a) Group-16

(b) Lithium

(c) Group-17

13. M.O. electronic configuration of  $\text{O}_2^-$

$$= \text{Bond order} = \frac{1}{2}(8 - 5) = \frac{3}{2} = 1.5$$

M.O. electronic configuration of  $\text{N}_2^+$

$$\text{Bond order} = \frac{1}{2}(7 - 2) = \frac{5}{2} = 2.5$$



As bond order of  $N_2^+ >$  bond order of  $O_2^-$ . Therefore,  $N_2^+$  is more stable than  $O_2^-$ .

14. (i) Due to strong hydrogen bonding or greater intermolecular hydrogen bonding in  $H_2O$  than in  $H-F$ .  $H_2O$  forms four bond with other water molecule wheres  $HF$  forms only two H-bonds.  
 (ii) Density of ice is less than (water) liquid form.  
 (iii) Due to H-bonding in  $HF$   $H-F \cdots H-F \cdots H-F$ . This can dissociate to give  $HF_2^-$  ion and hence  $KHF_2$  exists but no H-bnding in  $H-Cl$ . So  $HCl_2^-$  ion not exist.
15. (i)  $sp^3$ , pyramidal.  
 (ii)  $sp^3d$ , linear.  
 (iii)  $sp^2d^2$ , square planar.
16. (a) At  $-273^\circ C$ , volume of the gas becomes equal to zero *i.e.*, the gas ceases to exist. 1

(b) According to Charle's law :  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$$V_1 = 20 \text{ mL } V_2 = ? \text{ mL}$$

$$T_1 = 15 + 273 = 288 \text{ K} \quad T_2 = 35 + 272 = 308 \text{ K}$$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{20}{288} \times 308 = 21.38$$

Volume of hydrogen gas at  $35^\circ C = 21.38 \text{ ML}$

17. (a)  $\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$   
 (b) Greater the value of 'a', more easily the gas is liquefiable. Similarly, greater the value 'b' greater is the molecular size. Hence, gas 'y' will be more easily liquefiable and will have greater molecular size.
18. (a)  $CaCl_2, MgCl_2$   
 (b) In cation exchange resin  

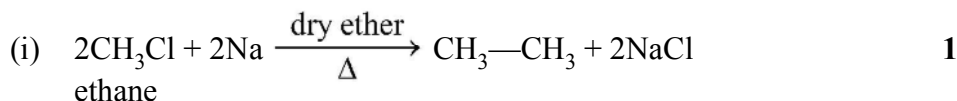
$$2RH_{(s)} + M_{(aq)}^{2+} \rightleftharpoons MR_{2(s)} + 2H_{(aq)}^+$$
  
 In anion exchange resis  

$$RNH_3^+ OH^- + X_{(aq)}^- \rightleftharpoons RNH_3^+ X_{(s)}^- + OH^-$$
  
 Finally  $H^+ + OH^- \rightleftharpoons H_2O_{(l)}$
19. (i)  $B(OH)_3 + 2H_2O \rightarrow B(OH)_4 + H_3O^+$  1  
 (ii)  $2Al + 2NaOH + 6H_2O \rightarrow 2Na[Al(OH)_4] + 3H_2$  1  
 (iii)  $BF_3 + NH_3 \rightarrow [H_3N \rightarrow BF_3]$  (Adduct) 1

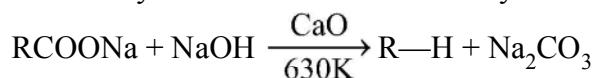
20. IUPAC Name :

- (i) Propylbenzene
- (ii) 2, 5 - Dinethyl heptane
- (iii) 3-Chloropropanal

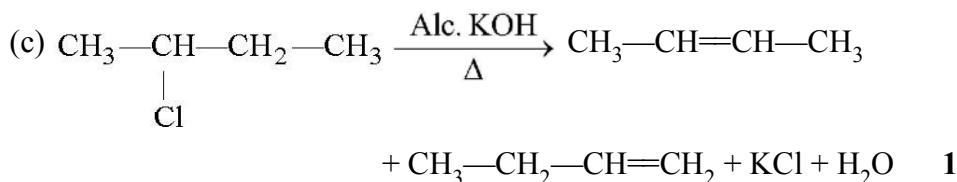
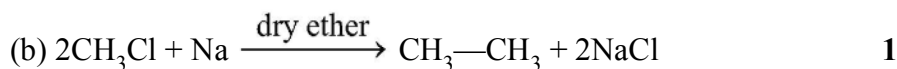
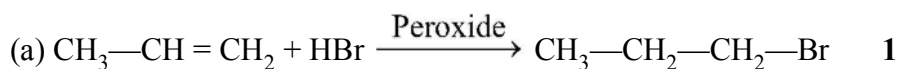
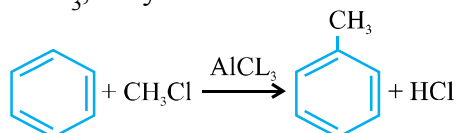
21. Wurtz's reaction :



(ii) Decarboxylation of sodium salt of fatty acid with soda line. 1



(iii) *Friedal Craft reaction* : When benzene treated with alkyl halide in presence of  $\text{AlCl}_3$ , Alkyl benzene is formed. 1

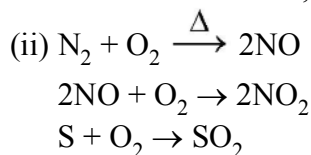


22. (a) (i) Like Al, Be is not readily attacked by acids because of the presence of an oxide film on the surface metal.

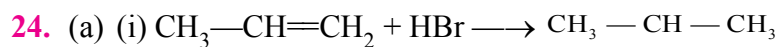
(ii)  $\text{Be}(\text{OH})_2$  dissolved in excess of  $\text{NaOH}$  to give  $[\text{Be}(\text{OH})_4]^{2-}$  just as  $\text{Al}(\text{OH})_3$  dissolve in  $\text{NaOH}$  to form  $[\text{Al}(\text{OH})_4]^{-1}$ .

(b)  $\text{Na}^+$  ions are present in blood plasma and in the interstitial fluid. These ions participate in the transmission of nerve signals and in regulating the flow of water across cell membranes.

23. (i) The symptoms of the villagers show that oxides of nitrogen and sulphur are coming out of the chimney of factory. This is due to combustion of fossil fuel like coal, gasoline etc.

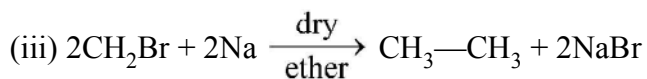
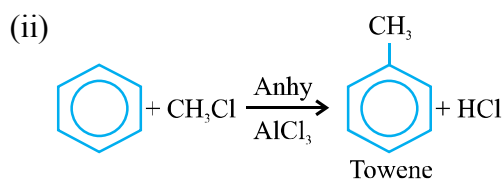


(iii) Environment concern, scientific attitude.

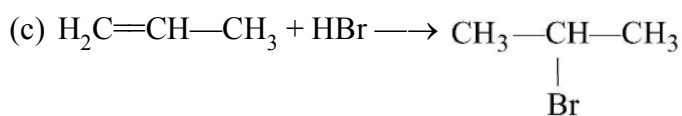
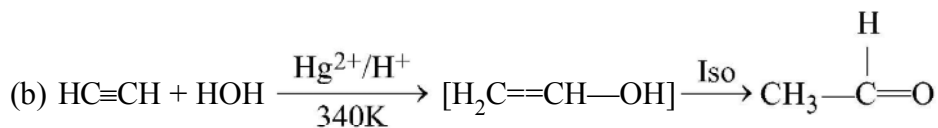
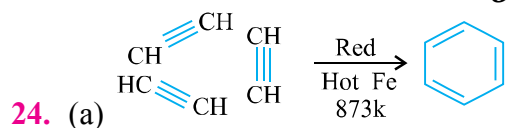
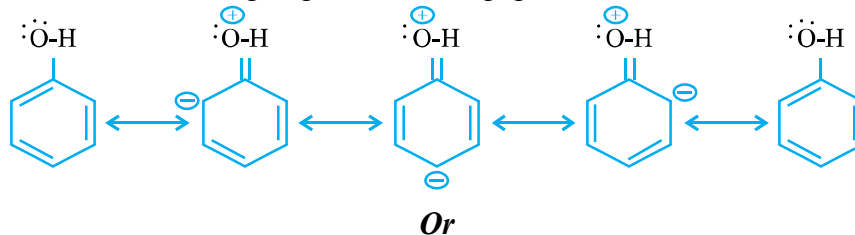


Br

2 - Bromopropane



- (b) This is because of + R (resonance effect) —OH group. The electron density is increased at *o*- and *p*-positions as compared to at *m*-position. Hence the new group will enter at *p*-position.



25. (i)  $K = \frac{1}{10} = 0.1$

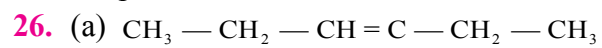


(iii) Total  $[\text{H}_3\text{O}^+] = [\text{H}_3\text{O}^+]_{\text{acid}} + [\text{H}_3\text{O}^+]_{\text{H}_2\text{O}}$   
 $= 10^{-8} + 10^{-7}$   
 $= 11 \times 10^{-8} \text{ M}$   
 $\text{pH} = -\log(\text{H}_3\text{O}^+)$   
 $\text{pH} = -\log(11 \times 10^{-8})$

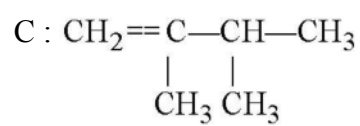
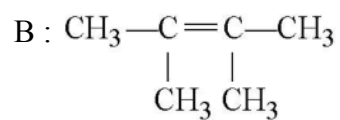
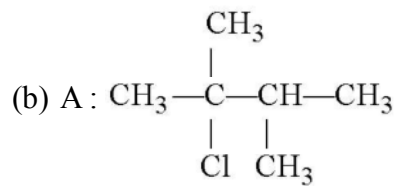
$$\text{pH} = -(\log 11 + \log 10^{-8})$$

$$\text{pH} = 8 - \log 11$$

$$\text{pH} = 8 - 1.0414 = 6.9586$$



3-Ethylhe - 3- one



**Common Annual School Examination, 2015-16**

**Subject : Chemistry**

**Class : XI**

**Time : 3 Hrs.]**

**[M.M. : 70**

- (i) All questions are compulsory.
- (ii) Q. No. 1 to 5 are Very Short Answer Question carrying 1 mark each.
- (iii) Q. No. 6 to 10 are Short Answer Questions and carrying 2 marks each.
- (iv) Q. No. 11 to 22 are Short Answer Questions and carrying 3 marks each.
- (v) Q. No. 23 is Value Based Questions carries 4 marks.
- (vi) Q. No. 24 to 26 are Long Answer Questions and carrying 5 marks each.
- (vii) Use log tables, if necessary, Use of calculator is not allowed.

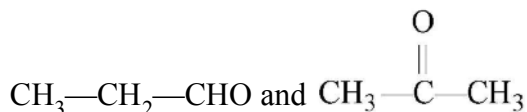
1. In a reaction  $A + B_2 \rightarrow AB_2$ , identify the limiting reagent when 2 mole of A are mixed with 3 mole of  $B_2$ .
2. Write the general electronic configuration of *f*-block elements.
3. Define critical temperature.
4. Classify the following as Lewis acid or Lewis base :  
 $NH_4^+$  and  $NH_3$
5. What is the oxidation number of Mn in  $KMnO_4$  ?
6. (a) How many sub-shells are associated with  $n = 4$  ?  
(b) How many electrons will be present in these sub-shell having  $m_s$  value of  $-\frac{1}{2}$  for  $n = 4$  ?
7. Give one point to differentiate the following thermodynamic terms :  
(a) Extensive properties and intensive properties.  
(b) Isothermal process and isobaric process.
8. Account for the following :  
(a)  $K_2CO_3$  cannot be prepared by Solvay process.  
(b) Alkali metals are not found in nature in free state.

**Or**

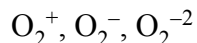
Write balanced equations for the equation between :

- (a)  $Na_2O_2$  and water
  - (b)  $Na_2O$  and  $CO_2$
9. Give suitable reasons for the following :  
(a)  $[SiF_6]^{2-}$  is known whereas  $[SiCl_6]^{2-}$  not.

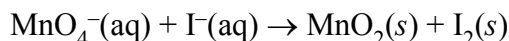
- (b) Diamond is covalent, yet it has high melting point.
10. (a) What type of isomerism is shown by pentane and 2-methyl butane ?  
 (b) Write the name of isomerism among the following compounds :



11. Calculate the concentration of nitric acid in moles per litre in a sample which has density  $1.40 \text{ g ml}^{-1}$  and the mass percent of nitric acid in it being 69%.  
 Molar mass of  $\text{HNO}_3 = 63 \text{ g mol}^{-1}$ .
12. Account for the following :  
 (a) An anion is always bigger than its parent atom.  
 (b) Chlorine (Cl) have more negative electron gain enthalpy than fluorine (F). [Atomic no. F = 9, Cl = 17]
13. Give the shapes of following covalent molecules using VSEPR theory :  
 (a)  $\text{ClF}_3$  (b)  $\text{XeF}_4$  (c)  $\text{AsF}_5$
14. Compare the relative stability of the following species on the basis of molecular orbital theory and indicate their magnetic properties :



15. (a) In terms of Charles's law, explain why  $-273^\circ\text{C}$  is the lowest temperature ?  
 (b) Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of  $1 \text{ dm}^3$  at  $27^\circ\text{C}$ .  
 [R =  $0.083 \text{ bar dm}^3 \text{ K}^{-1} \text{ mol}^{-1}$ ]
16. (a) For the reaction,  $2\text{Cl(g)} \rightarrow \text{Cl}_2\text{(g)}$ , what are the signs of  $\Delta H$  and  $\Delta S$  ?  
 (b) For the reaction at 298 K,  $2\text{A} + \text{B} \rightarrow \text{C}$ ,  $\Delta H = 400 \text{ kJ mol}^{-1}$  and  $\Delta S = 0.2 \text{ kJ K}^{-1} \text{ mol}^{-1}$ . At what temperature will the reaction become spontaneous ?
17. Balance the following redox reaction in basic medium : (Write steps of any one method used)

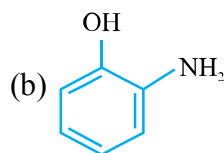


18. (a) Name the different salts that cause permanent hardness of water. (Any two)  
 (b) How do we obtain demineralised water from hard water after passing it from synthetic ion exchange resins ? Give reactions ?
19. (a) Mention two similarities in the behaviour of Be and Al to show that they have diagonal relationship.  
 (b) What is the biological importance of Na in our body ?

20. What happens when :

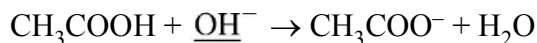
- (a) Boric acid is added to water.
- (b) Al is treated with dil. NaOH.
- (c)  $\text{BF}_3$  is treated with  $\text{NH}_3$ .

21. Write the IUPAC names of the following :



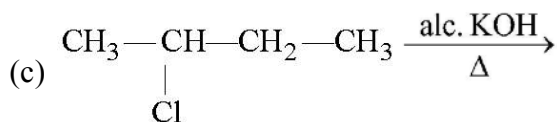
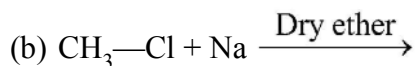
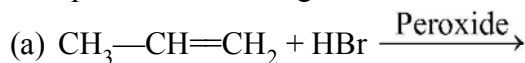
*Or*

(a) Identify the reagent shown underlined as electrophile or nucleophile :



(b) On complete combustion of 0.246 g of an organic compound gave 0.198 g of  $\text{CO}_2$  and 0.1014 g of  $\text{H}_2\text{O}$ . Determine the percentage composition of carbon and hydrogen in the compound.

22. Complete the following reactions :



23. A factory was started near a village. Suddenly villagers started feeling the presence of irritating vapours in the village and cases of headache, chest pain, cough, dryness of throat and breathing problems increased. Anil a science student blamed the emissions from the chimney of the factor for such problems.

After reading the above paragraph, answer the following questions :

(a) Explain what could have happened ?

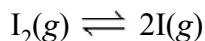
- (b) Give chemical reactions for the support of your explanation.
- (c) Mention the values shown by the student in the above paragraph.
24. (a) Write the electronic configuration of Cr( $Z = 24$ ). Why is it different from the expected configuration ?
- (b) The mass of an electron is  $9.1 \times 10^{-31}$  kg and its kinetic energy is  $3 \times 10^{-25}$  J. Calculate its wavelength.
- (c) Which of the following orbitals is not possible and why ?  
 $2d, 2s, 3p, 3s$

**Or**

- (a) Calculate the wavelength and frequency of limiting line of Lyman series (Rydberg constant =  $109677 \text{ cm}^{-1}$ ).
- (b) Give quantum numbers for electrons with highest energy in sodium atom ( $Z = 11$ ).
- (c) Which of the following sets of quantum number are not possible ? Give reasons :
- (i)  $n = 1, l = 0, m_l = 0, m_s = -\frac{1}{2}$
- (ii)  $n = 0, l = 0, m_l = 0, m_s = -\frac{1}{2}$
25. (a) Equilibrium constant for a reaction is 10. What will be the equilibrium constant for the reverse reaction ?
- (b) Write the conjugate acids for the Bronsted base  $\text{OH}^-$  and  $\text{CH}_3\text{COO}^-$ .
- (c) Determine the pH of  $10^{-8}$  M HCl solution taking into account the  $\text{H}^+$  produced by water also. ( $\log 11 = 1.0414$ )

**Or**

- (a) At certain temperature and total pressure of  $10^5$  Pa, iodine vapour contains 40% by volume of I atoms :



Calculate  $K_p$  for the equilibrium.

- (b) What is the effect of :



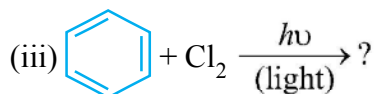
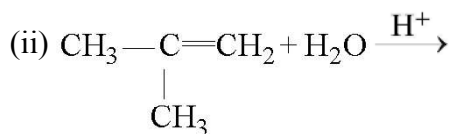
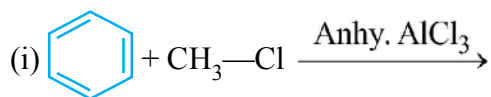
- (i) Addition of  $H_2$
- (ii) Removal of CO

on the equilibrium :  $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

(c) Mention one application of solubility product.

**26.** (a) Propanal and pentan-3-one are the ozonolysis product of an alkene. What is the structural formula of the alkene ?

(b) Give the main products of the reactions :



**Or**

(a) An alkyl halide (A) of formula  $C_6H_{13}Cl$  on treatment with alcoholic KOH give two isomeric alkenes (B) and (C) ( $C_6H_{12}$ ). Both alkenes on hydrogenation give 2, 3-Dimethyl butane. Predict the structure of A, B and C.

(b) Why does benzene show electrophilic substitution easily ?

(c) Name the compound that will be required to obtain butane using Kolbe's electrolysis process.