11th STD - CHEMISTRY SPECIAL GUIDE

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XI - CHEMISTRY SPECIAL GUIDE

UNIT 1. BASIC CONCEPTS AND CHEMICAL CALCULATIONS

2 & 3 Mark questions

1. Define Relative atomic mass.

- Relative atomic mass $= \frac{\text{Average mass of the atom}}{\text{Unified atomic mass}}$

2. Define Mole.

- One mole is the amount of substance that contains as many elementary particles present in 12g of carbon -12 isotope.

3. Define Avogadro number.

- The number of atoms present in 1 mole of substance.
- It's value is 6.023×10^{23}

4. Define Molar mass.

- The mass of 1 mole of a substance.

5. Define Molar volume.

- The volume occupied by one mole of a gas at a given temperature and pressure.

6. Define Gram equivalent mass.

- The mass of the substance that combines or displaces 1.008g of hydrogen or 8g of oxygen or 35.5g of chlorine.

7. Define Basicity.

- The number of moles of ionizable H^+ ions present in one mole of an acid.

8. Define Acidity.

- The number of moles of ionizable of OH^- ions present in one mole of a base.

9.Define Oxidation number.

- It is the imaginary charge left on the atom when all the other atoms have been removed with their usual oxidation States assigned by a set of rules.

10. Give the difference between Oxidation and Reduction.

S.No	Oxidation	Reduction
1.	Addition of Oxygen	Removal of Oxygen
2.	Removal of Hydrogen	Addition of Hydrogen
3.	Removal of Electrons	Addition of Electrons

4.	Increase in Oxidation number	Decrease in Oxidation number

11. What is Limiting and excess agents.

- The reagents which limits the further reaction to take place is called as limiting agents.
- The reagent which is present in excess is called as excess agents.

12. What is the Empirical formula of the following?

i) Fructose ($C_6H_{12}O_6$)

ii) Caffeine ($C_8H_{10}N_4O_2$)

Compound	Empirical formula
Fructose	CH_2O
Caffeine	$C_4H_5N_2O$

13. Calculate the gram equivalent mass of H_2SO_4.

- Gram equivalent mass = $\frac{\text{molar mass}}{\text{Basicity}} = \frac{98}{2}$ = 49 g equ $^{-1}$

14. Write the difference between molar mass and molecular mass. Calculate the molar mass and molecular mass for Carbon monoxide.

S.No	Molecular mass	Molar mass
1	Ratio of the mass of a molecule to the unified atomic mass	The mass of one mole of the substance
2	Unit – amu	Unit - g mol-1
3	Molecular mass of CO is 28 amu	Molar mass of CO is 28 g.mol $^{-1}$

UNIT 2. QUANTUM MECHANICS MODEL OF ATOM

2 & 3 Mark questions

1. Describe the Aufbau principle.

- In orbitals, the electrons are filled in the order of their increasing energy.

2. State and explain Pauli exclusion principle.

- No two electrons in an orbital can have same set of values of all four quantum numbers.

3. Define Hund rule.

- Electron pairing is not possible until all the orbitals contain one electron each.

4. What are the limitation of Bohr atomic model theory.

- This model is valid only for one electron species like Hydrogen. It fails for multi electron atoms.
- It fails to explain Zeeman effect and Stark effect.

5. Define Heisenberg's Uncertainty Principle.

- It is impossible to determine both the position and momentum of microscopic particle simultaneously and accurately.
- $\Delta x . \Delta p \geq h/4\pi$
- Δx - uncertainty in the position
- Δp - uncertainty in the momentum

6. How many orbitals are possible n=4?

- One s - orbital
- Three p – orbitals
- Five d - orbitals
- Seven f - orbitals
- Total 16 orbitals

7. Define Exchange energy.

- If two electrons of same spin are present in an orbital, they may exchange their positions.
- During the exchange, energy is released. This released energy is called as exchange energy.

8. Write the electronic configuration of Mn^{2+} and Cr^{3+}

- $Mn^{2+} = [Ar]3d^5$
- $Cr^{3+} = [Ar]3d^3$

9. Write the electronic configuration of Cu and Cr.

- $Cu = [Ar]3d^{10}\ 4s^1$
- $Cr = [Ar]3d^5\ 4s^1$

10. For each of the following give the sub level designation, the allowable m values and the number of orbitals. 1.n=7,l=0 ,. 2.n=3,l=1 3.n=4,l=2 4.n=5,l=3

S.No	n	l	Sub energy levels	m values	Number of orbitals
1	7	0	7s	0	1
2	3	1	3p	-1,0,+1	3
3	4	2	4d	-2,-1,0,+1,+2	5
4	5	3	5f	-3-2,-1,0,+1,+2,+3	7

11. How many radial nodes, and angular nodes for 2s,4p,5d, 4f orbitals?

orbital	n	l	Radial node (n-l-1)	Angular node (l)
2s	2	0	1	0
4p	4	1	2	1
5d	5	2	2	2
4f	4	3	0	3

5 Mark questions

1. Explain the Bohr atomic model theory.

- The energy of an electron is Quantized.
- The electrons are moving around the nucleus in circular path called as stationary orbits.
- The angular momentum of the electron is given by $mvr = nh/ 2\pi$
- As long as an electron is in a fixed orbit, it will not lose energy. But if the electron jumps from higher energy level to lower energy level it will emit radiation.
- If energy is supplied, the electron will jump from lower energy level to higher energy level.

2. Derive the De-Broglie equation of wave nature of electrons.

- $E = mc^2$
- $E = hv$
- $hv = mc^2$
- $v = c / \lambda$
- $hc / \lambda = mc^2$
- $\lambda = h/mc$
- $\lambda = h/mv$

3. Explain Quantum numbers and its types.

a) Principal Quantum Numbers (n) :

- It represents energy level.
- The maximum number of electrons in a shell can be calculated by using formula $2n^2$.
- $E = (-1312.8) Z^2/ n^2$ KJ/ mol^{-1}
- $r = (0.529\ n^2) / Z$ A°

Shell	K	L	M	N
n- value	1	2	3	4

b) Azimuthal or Subsidiary Quantum Number (l):

- It represents sub shell.
- The maximum number of electrons in an orbital is given by 2(2l+1).
- Angular momentum = $\sqrt{l(l+1)}\ \frac{h}{4\pi}$

Orbital	s	p	d	f
l – value	0	1	2	3

c) Magnetic Quantum Number (m) :

- It represents different orientation of orbitals in space.
- Its value ranges from - l to + l through zero.
- The angular momentum is given by the azimuthal quantum number and the direction is given by the magnetic quantum number.

d) Spin Quantum Number (ms):

- It represents the spin of the electron.
- The electron in an atom revolves around the nucleus and also spins in clockwise or in the anticlockwise direction.

Spin direction	Clockwise	Anti clockwise
s – value	+1/2	-1/2

UNIT 3. PERIODIC CLASSIFICATION OF ELEMENTS

2 & 3 Mark questions

1. Define Modern periodic law.

- The physical and chemical properties of the elements are the periodic functions of the atomic numbers.

2. Define Shielding effect or Screening effect.

- The inner shell electrons act as a shield between the nucleus and the valence electrons, this is called shielding effect.

3. Define Electronegativity.

- Electronegativity is a tendency of an element present in covalent molecules to attract the shared pair of electrons towards itself.

4. Define Diagonal relationship.

- The similarities in the properties between the diagonally present elements are called as diagonal relationship.
- Ex. Li and Mg.

5. Why Noble gases (Neon) has zero electron affinity?

- Neon has Stable Fully filled ns^2np^6 electronic configuration. So it will not accept electrons.
- $Ne_{10} = 1s^2 2s^2 2p^6$

6. Define Isoelectronic ions.

- Ions having same number of electrons and same electronic configuration are called as Isoelectronic ions.
- Ex. $Na^+ = 2, 8$
  $F^- = 2, 8$

7. Define Effective nuclear charge.

- The net nuclear charge experienced by the valence electron in the outermost shell is called as effective nuclear charge.
- $Z_{eff} = Z - S$

8. Why Halogens act as oxidizing agents?

- Halogens have high electro negativity and electron affinity.
- Halogens have unstable np^5 electronic configuration.
- Hence it accepts one electron and become a stable Noble gas configuration.

9. Give any two anomalous properties of 2nd group elements.

- Lithium and Beryllium form covalent compounds but other forms ionic compounds.
- The second period has only 4 orbitals in the valence shell. The maximum valency is 4. but others have higher valences.

5 Mark questions

1. Explain the pauling method for the determination of ionic radius.

- $d = rC^+ + rA^-$
- d = internuclear distance between the ions
- rC^+ =Radius of cation
- rA^- = Radius of anion
- $Z_{eff} = Z - S$
- $\frac{rc+}{rA} = \frac{Z*(A-)}{Z*(c+)}$
- rC^+ and rA^- can be calculated from the above equation.

UNIT 4. HYDROGEN

2 & 3 Mark questions

1. What are the similarity of Hydrogen with Alkali metals.

- Hydrogen have $1s^1$ and Alkali metals have ns^1 electronic configuration.

- Like Alkali metals, Hydrogen forms unipositive ions. Ex. H^+ and Na^+.
- It also acts as a reducing agent.

2. How Hydrogen differs from Halogens.

- Hydrogen has less electron affinity than halogens.
- The tendency of forming Hydride ions is low.
- But halogens easily form Halides.

3. Define Isotopes and explain the three types of isotopes of Hydrogen.

- Elements having same atomic number but different mass number is called as isotopes.
- Hydrogen has 3 isotopes.
- Protium - $_1H^1$
- Deuterium - $_1H^2$
- Tritium - $_1H^3$

4. Explain Ortho and Para Hydrogen.

- In hydrogen molecule, if two nuclei rotates in the same direction is called as Ortho Hydrogen.
- In hydrogen molecule, if two nuclei rotates in the opposite direction is called as Para Hydrogen.

5. How will you convert Para Hydrogen to Ortho Hydrogen.

- By using catalyst like Iron.
- By passing electric discharge.
- By heating at 800°C.
- By mixing with paramagnetic molecules like oxygen.
- By mixing with atomic hydrogen.

6. Explain the water gas shift reaction.

- $CO + H_2O \xrightarrow{400^oC} H_2 + CO_2$

7. Discuss the three types of covalent hydrides.

- Electron deficient - B_2H_6
- Electron rich hydrides - Water
- Electron precise - Methane

8. List the uses of Heavy water.

- It is used as Moderators in nuclear reactor.
- It is used as tracer element to study the mechanisms of organic reactions.
- It is used as coolant in nuclear reactors.

9. List the uses of Deuterium.

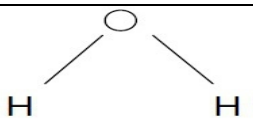
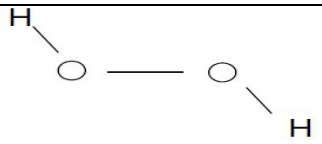
- It is used to prepare heavy water.

- It is used as tracer element to study the mechanisms of organic reactions.
- High speed deuterium is used in artificial radio activity.

10. Explain the exchange reaction of Deuterium.

- $CH_4 + 2D_2 \longrightarrow CD_4 + 2H_2$

11. Compare the structures of H_2O and H_2O_2.

S.No	H_2O	H_2O_2
1.	It is bent structure.	It is open book structure
2.	The bond angle is 104.5°	The bond angle is 90.2°
3.		

12. NH_3, H_2O, and HF arrange in the order of hydrogen bonding.

- $HF > H_2O > NH_3$

13. Differentiate Ortho Hydrogen and Para Hydrogen.

S.No	Ortho Hydrogen	Para Hydrogen
1	The spins of two Hydrogen nuclei exist in same direction	The spins of two Hydrogen nuclei exist in opposite direction
2	75 % at room temperature	25 % at room temperature
3	More stable	less stable

UNIT 5. ALKALI AND ALKALINE EARTH METALS

2 & 3 Mark questions

1. What are the general characteristics of Alkali metals.

- Their general electronic configuration is ns^1.
- Their common oxidation state is +1.
- Moving down the group the ionic radius increases.

2. What is the reason for distinctive behaviour of Lithium or Beryllium.

- Smaller in size.
- High polarizing power and High hydration energy.
- Absence of d-orbitals.

3. What are the similarities (diagonal relation) between Lithium and Magnesium.

- Both are Hard.
- Both reacts slowly with water.

- Both do not form Super oxides.

4. Explain the important common features of group 2 elements.

- Their general electronic configuration is ns^2.
- Their common oxidation state is +2.
- Moving down the group the ionic radius increases.

Beryllium carbides give methane on hydrolysis.

5. Explain the diagonal relation (similarities) between Beryllium and Aluminium.

- Both renders passive with Nitric acid.
- Both hydroxides are Amphoteric in nature.
- Both carbides give methane on hydrolysis.

6. List the uses of Gypsum.

- It is used to make Plaster of Paris.
- It is used in tooth paste, Shampoos and hair products.
- It is used to make Portland cement.

7. List the uses of Plaster of Paris.

- It is used in the building industry
- It is used in the treatment of bone fracture.
- It is used for making statues and casts.

8. How is plaster of paris prepared?

- $2\ CaSO_4.2H_2O \xrightarrow{393\ K} 2\ CaSO_4.\frac{1}{2}H_2O + 3\ H_2O$

9. Describe briefly the biological importance of sodium and Potassium.

- Found in biological fluids.
- Maintain the ion balance and nerve impulse conduction.

10. Describe briefly the biological importance Calcium and Magnesium.

- Essential for DNA synthesis.
- Balancing the electrolytes in our body.

UNIT 6. GASEOUS STATE

2 & 3 Mark questions

1. Define Boyle's law.

- At constant temperature the volume of a gas is inversely proportional to the pressure.
- $V \propto 1/P$

2. Define Charles law.

- At constant pressure the volume of a gas is directly proportional to the temperature.
- $V \propto T$

3. State Gay-Lussac's law.

- At constant volume the pressure of a gas is directly proportional to the temperature.
- $P \propto T$

4. Define Avogadro Hypothesis.

- Equal volume of all gases under same temperature and pressure contains equal number of molecules.
- $V \propto n$

5. Define Dalton law of partial pressure.

- The total pressure of a gaseous mixture of nonreacting gases is equal to the sum of the partial pressure of the gases present in the mixture.

- $P_{TOTAL} = P_1+P_2+P_3+\text{------}$

6. Define Graham's law of diffusion.

- The Rate of diffusion of a gas is inversely proportional to the square root of the molar mass.
- Rate of diffusion $\propto \frac{1}{\sqrt{M}}$

7. Define Critical temperature.

- The temperature above which a gas cannot be liquefied even at a high pressure.

8. Define Critical pressure.

- The pressure required to liquefy one mole of a gas at its critical temperature.

9. Define Critical volume.

- The volume occupied by one mole of as gas at its critical temperature and pressure.

10. Define Joules Thomson effect.

- The lowering of temperature, when a gas is made to expand adiabatically from high pressure to low pressure is called as Joules Thomson effect.

11. What is Inversion temperature.

- The temperature below which a gas obeys Joule Thomson effect is called as Inversion temperature.
- $T_i = 2a/Rb$

12. Difference between Diffusion and Effusion.

Diffusion	Effusion
The movement of the gas molecules through another gas from high concentration to low concentration is called as Diffusion.	The movement of the gas molecules through a small hole from high concentration to low concentration is called as Effusion.

13. What are ideal gases? Give the difference between Ideal and real gases.

- Gases which obey PV = nRT are called Ideal gases.

S.No	Ideal Gas	Real Gas
1.	Ideal gases obey Ideal gas equation under all conditions of temperature and pressure	Real gases obey the Ideal gas equation at low pressure and high temperature
2.	There is no force of attraction	There is a force of attraction
3.	They obey Ideal gas equation PV=nRT	They obey the Vanderwaals equation.

14. Derive the Ideal gas equation.

- Boyle's law - $V \propto 1/P$
- Charles law - $V \propto T$
- Avogadro's law - $V \propto n$
- $V \propto nT/P$
- $PV = nRT$.

15. Derive the critical constants from vander Waals equation constants.

- $(P+an^2/V^2)(V-nb) = nRT$
- $Vc = 3b$
- $Pc = a/ 27b^2$
- $Tc = 8a/ 27Rb$.

16. What is compressibility factor?

- The deviation of real gases from ideal behaviour is measured by the ratio of PV to nRT.
- $Z = PV/nRT$

17. What are the different methods used for liquefaction of gases?

- Linde's method
- Claude's method
- Adiabatic process

UNIT 7. THERMODYNAMICS

2 & 3 Mark questions

1. Define Extensive property.

- The properties that depends on mass is called as Extensive property.
- Ex. Volume and Mass

2. Define Intensive property.

- The property that does not depend on mass is called as Intensive property.
- Ex. Density and temperature

3. Define Reversible process.

- The process in which the system can be restored to the initial state from the final state is called as Reversible process.

4. Define Irreversible process.

- The process in which the system cannot be restored to the initial state from the final state is called as Irreversible process.

5. Define Adiabatic process.

- There is no exchange of energy, between the system and the surrounding is called as Adiabatic process.
- $q = 0$

6. Define Isothermal process.

- The temperature remains constant, when the system moves from initial to final state is called as Isothermal process.
- $dT = 0$

7. Define Isobaric process.

- The Pressure remains constant, when the system moves from initial to final state is called as Isobaric process.
- $dP = 0$

8. Define Isochoric process.

- The Volume remains constant, when the system moves from initial to final state is called as Isochoric process.
- $dV = 0$

9. Define State function.

- Thermodynamic property of a system whose value does not depends on the path by which the system changes from its initial to final state.
- Ex. P, V and T

10. Define Path function.

- Thermodynamic property of a system whose value depends on the path by which the system changes from its initial to final state.
- Ex. Heat and Work.

11. Define Zeroth law of thermodynamics.

- If two systems are in thermal equilibrium with a third one, then they tend to be thermal equilibrium with themselves.

12. Define First law of thermodynamics.

- Energy can neither be created nor be destroyed, but one form of energy can be converted in to another form.

13. Define Third law of thermodynamics.

- The entropy of a pure crystalline substance at absolute zero is zero.

14. Define Specific heat capacity.

- The amount of heat absorbed by one Kg of a substance to raise its temperature by one Kelvin is called as Specific heat capacity.

15. Define Molar heat capacity.

- The amount of heat absorbed by one mole of a substance to raise its temperature by one Kelvin is called as Molar heat capacity.

16. Define the Calorific value of food. Write its unit?

- The amount of heat produced when one gram of a substance is completely burnt, is called as calorific value of food.
- Unit. JKg^{-1}

17. Define Exothermic and Endothermic reactions.

Exothermic Reaction	Endothermic Reaction
When Energy is liberated by the system to the surrounding is called as Exothermic Reaction	When Energy is absorbed by the system from the surrounding called as Endothermic Reaction.

18. Define Hess's law.

- The enthalpy change of a reaction at constant volume or pressure is same whether it takes place in single or multiple steps provided the initial and final states are same.

19. Define Lattice energy.

- The amount of energy required to completely remove the ions from its crystal lattice to an infinite distance is called as lattice energy.

20. Define Gibb's free energy.

- G=H-TS
- G - Gibbs free energy
- H - Enthalpy
- S – Entropy
- T- Temperature

21. Define spontaneous process and the condition.

- The reaction which takes place without any external driving force is called as spontaneous process.

Conditions for Spontaneity:

- ΔS = +ve
- ΔH = -ve
- ΔG = -ve

5 Mark questions

1. What are the various statements for Second law of thermodynamics.

- **Kelvin Planck statement**
  It is impossible to construct a machine that absorbs heat from a hot source and converts it completely into work by a cyclic process without transferring a part of heat to a cold sink.

- **Clausius Statement**
  It is impossible to transfer heat from a cold sink to a hot sink without doing some work.

- **Entropy statement**
  For a spontaneous process, the entropy of an isolated system will increase.

- **Efficiency statement**
  The efficiency of a machine never be 100%

- **Efficiency**

$$\text{Efficiency} = \frac{\text{work performed}}{\text{heat absorbed}} \times 100\%$$

2. Give the characteristics of Internal energy.

- Extensive property.
- State function.
- For a cyclic process, $\Delta U = 0$.
- The change in the internal energy is $\Delta U = U_f - U_i$
- If $U_f > U_i$ then ΔU = + ve

  If $U_f < U_i$ then ΔU = -ve.

3. What are the characteristics of Gibb's free energy.

- $G = H - TS$
- Extensive property.
- State function.
- Spontaneous process ΔG = -ve
  Nonspontaneous process ΔG = +ve
  Equilibrium process $\Delta G = 0$.
- Net Work done $(-\Delta G) = -W - P\Delta V$.

UNIT 8. PHYSICAL AND CHEMICAL EQUILIBRIUM

2 & 3 Mark questions

1. Why chemical equilibrium is considered as Dynamic equilibrium.

- At equilibrium the forward and the backward reactions will proceed at the same rate.
- No macroscopic changes is observed.

2. Define Homogenous and Heterogeneous equilibrium.

- When the reactants and the products are in the same phase it is called as Homogeneous equilibrium.
- When the reactants and the products are in the different phase it is called as Heterogeneous equilibrium.

3. Define Equilibrium constant.

- (At equilibrium conditions) $K_c = \frac{\text{Product of the active masses of the products}}{\text{Product of the active masses of the reactants}}$

4. Define Reaction quotient.

- (At non equilibrium conditions) $Q = \frac{\text{Product of the active masses of the products}}{\text{Product of the active masses of the reactants}}$

5. Define Law of mass action.

- At a given temperature, the rate of a chemical reaction is directly proportional to the product of the active masses of the reactants.
- Rate $\propto [\text{ Reactant }]^x$

6. Define Le - chatelier Braun principle.

- If a system at equilibrium is subjected to a disturbance, then the system will move in the direction to nullify the effect of the disturbance.

7. What is the relationship between Kp and Kc when Δng = 0, +ve,-ve.

- When $\Delta ng = 0$ $K_p = K_c$
- When $\Delta ng = +ve$ $K_p > K_c$
- When $\Delta ng = -ve$ $K_p < K_c$

5 Mark questions

1. Derive the relation between Kp and Kc.

- $aA + bB \Leftrightarrow cC + dD$
- $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

- $K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$

- $PV = nRT$

- $P = \frac{n}{V}RT$

- $K_p = \frac{[[C]\,RT]^c}{[[A]\,RT]^a} \times \frac{[[D]\,RT]^d}{[[B]\,RT]^b}$

- $K_p = \frac{[C]^c\,[D]^d}{[A]^a\,[B]^b} \times \frac{[RT]^{c+d}}{[RT]^{a+b}}$

- $K_p = K_c \times [RT]^{(c+d)-(a+b)}$

- $K_p = K_c \times [RT]^{\Delta ng}$

UNIT 9. SOLUTIONS

2 & 3 Mark questions

1.Define Molality.

- Molality = $\frac{\text{Number of moles of the solute}}{\text{Mass of the solvent in Kg}}$

2. Define Molarity.

- Molarity = $\frac{\text{Number of moles of the solute}}{\text{Volume of the solution in litre}}$

3. Define Normality.

- Normality = $\frac{\text{Number of gram equivalent of the solute}}{\text{Volume of the solution in litre}}$

4. Define Henry's law.

- The partial pressure of the gas in the vapour phase is directly proportional to the mole fraction of the solute at low concentration.
- $p_{solute} \alpha x_{Solute\ in\ solution.}$

5. Define Roult's law.

- The partial vapour pressure of the solution of volatile liquids of each component is directly proportional to its mole fraction.
- $p_A \alpha X_A.$

6. Define Colligative property.

- The property which depends on the number of solute particles is called as Colligative property.
- Ex. Osmotic pressure.

7. Define Elevation in the Boiling point.

- The increase in the Boiling point of a solvent when a solute is added, is called as Elevation in the Boiling point.

8. Define Ebullioscopic constant. (mole elevation constant).

- The elevation in the boiling of one molar solution is called as Ebullioscopic constant.

9. Define Depression in freezing point.

- The decrease in the Freezing point of a solvent when a solute is added, is called as Depression in the Freezing point.

10. Define Cryoscopic constant. (mole depression constant)

- The depression in the Freezing point of one molar solution is called as Cryoscopic constant.

11. Define Relative lowering of vapour pressure.

- The ratio between the lowering of vapour pressure and the vapour pressure of the pure solvent is called as Relative lowering of vapour pressure.

12. Define osmosis.

- The movement of the solvent molecules from lower concentration to higher concentration through a semi permeable membrane is called as Osmosis.

13. Define Osmotic pressure.

- The pressure applied on the solution to stop the movement solvent through a semi permeable membrane is called as Osmotic pressure.

14. What are Isotonic solutions.

- Solutions having same osmotic pressure at a given temperature are called as Isotonic solutions.

15. What are colligative properties?

- Relative lowering of vapour pressure
- Elevation of boiling point
- Depression in freezing point
- Osmotic pressure

16. What is Van't Hoff factor?

$$i = \frac{\text{Observed Colligative property}}{\text{Calculated Colligative property}}$$

17. What is Van't Hoff equation of Osmotic pressure?

- $\pi = CRT$.
- C - Concentration
- T - Temperature.
- R - Gas constant

UNIT 10. CHEMICAL BONDING

2 & 3 Mark questions

1. Define Bond order.

- Bond order is the number of Covalent Bonds between two atoms.
- $\text{Bond order} = \frac{Nb - Na}{2}$

2. Define Hybridization.

- The mixing of atomic orbitals of same atoms with same energy to give equivalent number of orbitals with same energy.

3. Define σ (Sigma) bond.

- When two atomic orbitals overlap linearly it forms Sigma bond.

4. Define π (Pi) bond.

- When two atomic orbitals overlap Sideways it forms Pi-bond.

5. Define Dipole moment.

- The polarity of a covalent bond can be measured by using the dipole moment. Its unit is Debye.
- Dipole moment = q x 2d

6. Explain Fajan's rule?

- When the charge of the cation or anion increases the covalent character also increases.
- The smaller cation and larger anion show greater covalent character.

UNIT 11. FUNDAMENTALS OF ORGANIC CHEMISTRY

2 & 3 Mark questions

1. What are the characteristic of Organic compounds.

- They are covalent compounds.
- They are insoluble in water but soluble in Organic solvents like Benzene.

- They are highly inflammable.
- They have low melting and boiling points.
- They form isomerism.

2. Explain the Homologous Series.

- They contain a characteristics functional group.
- Successive member differs by -CH_2 group in their molecular formula.
- They are prepared by similar methods.
- They have similar chemical properties.
- They have similar general formula.
- Ex. Alkane = C_nH_{2n+2}

3. What is the conditions for Optical isomerism.

- The molecule should have Chiral Carbon.

4. What is meant by functional group?

- An atom or group of atoms that react in characteristic way in an organic compound.
- Ex. -OH, -CHO.

5. What is lassaigne's extract or Sodium fusion extract?

- Dry sodium is melted in fusion tube and organic compound is added.
- Then heat the tube and broke in the distilled water, boiled and filtered.

6. Describe optical isomerism with example?

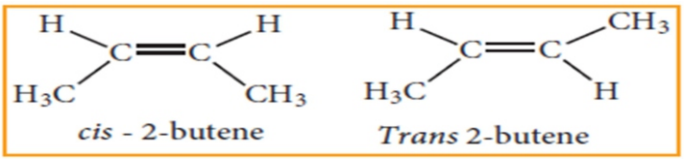
- Compounds having same property but different rotation of plane polarized light.
- Ex. lactic acid

5 Mark questions

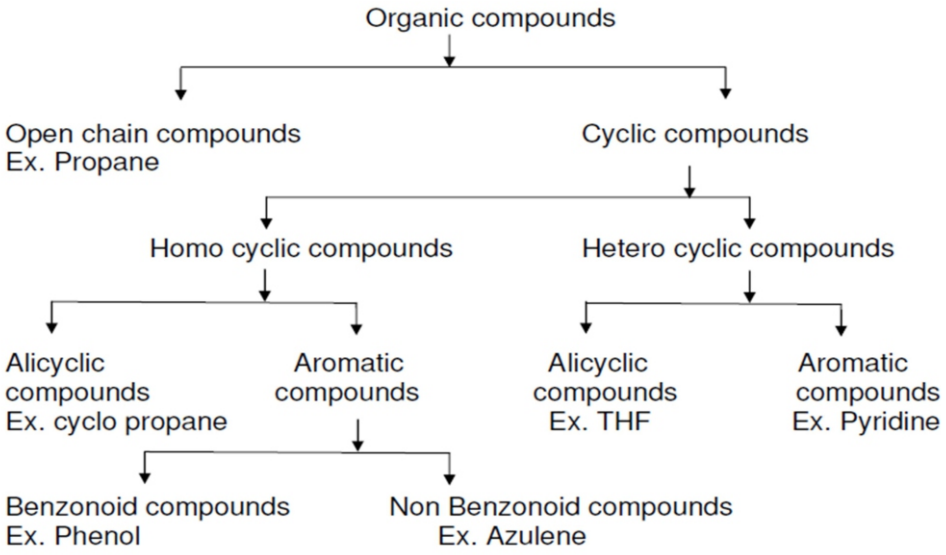
1. Explain the (Geometrical isomerism) or Cis - Trans isomerism using 2 - butane.

- **Cis isomers**
  Two similar groups are present on the same side of the carbon carbon double bond.

- **Trans isomers**
  Two similar groups are present on the opposite side of the carbon carbon double bond.



2. Explain the classification of organic compounds.



UNIT 12. BASIC CONCEPTS OF ORGANIC REACTIONS

2 & 3 Mark questions

1. Give the difference between Nucleophiles and Electrophiles.

Nucleophiles	Electrophiles
They are attracted towards to electron deficient centers. Ex. OH^-.	They are attracted towards to Electron rich centers. Ex. H^+.
Negative charged ions.	Positive charged ions.
Lewis bases.	Lewis acids.

2. Define Inductive effect.

- The change in the polarization of a covalent bond due to the presence of a adjacent atoms in a molecule.

3. Define Hyperconjugation.

- The delocalisation of electrons of σ-bond is called as Hyperconjugation.

4. Explain the Electromeric effect.

- Electro metric effect is a temporary effect in unsaturated compounds, in the presence of an attacking reagent.

5. Define Resonance or mesomeric effect.

- When an organic compound is represented by more than one structure, which differs only in the position of bonding and lone pair of electrons.

UNIT 13. HYDROCARBONS

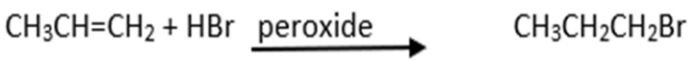
2 & 3 Mark questions

1. Explain Markovnikoff's rule with suitable example?

- Unsymmetrical alkene reacts with hydrogen halide, the hydrogen adds to the carbon having more number of hydrogen and halide adds to the carbon having less number of hydrogen.

2. Explain Peroxide effect or Kharash effect or Anti Markovikoff's rule?

- Unsymmetrical alkene reacts with hydrogen halide in presence of peroxides, the hydrogen adds to the carbon having less number of hydrogen and halide adds to the carbon having more number of hydrogen.



3. How does Huckel rule help to decide the aromatic character of aromatic compounds.

- The molecule must be co-planar.
- Complete delocalization of π electron in the ring.
- Presence of (4n + 2) π electrons in the ring.

4. Suggest a simple chemical test to distinguish propane and propene?

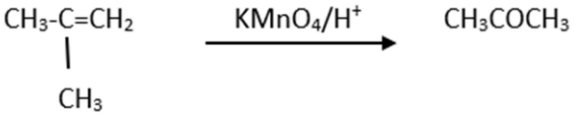
- Propene decolourises bromine water whereas propane does not.
- Propene decolourises acidified potassium permanganate whereas propane does not.

5. How will you distinguish 1-butyne and 2-butyne?

- 1-Butyne reacts with Tollen's reagent.
- 2-Butyne does not react with Tollen's reagent.

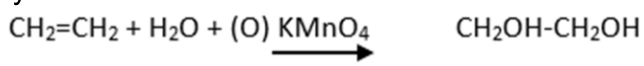
6. What happens when isobutylene is treated with acidified potassium permanganate?

- Acetone is formed.



7. What happens when ethylene is passed through cold dilute alkaline potassium permanganate?

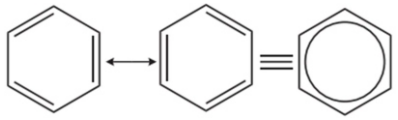
- Glycol is formed.



8. Explain the structure of benzene?

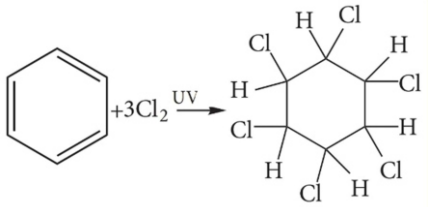
- Molecular formula C_6H_6.

- Benzene reacts with bromine in the presence of $AlCl_3$ to form mono bromobenzene, so benzene has a cyclic structure.
- Benzene consists of cyclic planar structure with alternate single and double bonds.
- Resonance structure of benzene.

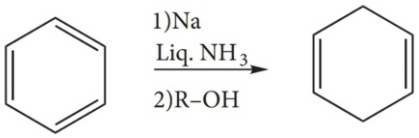


9. What is BHC? Write its preparation.

- BHC is benzene hexa chloride. It is known as gammaxane or Lindane.



10. Write Birch reduction?



UNIT 14. HALOALKANES AND HALOARENES

2 & 3 Mark questions

1. Compare SN^1 and SN^2 reaction mechanism.

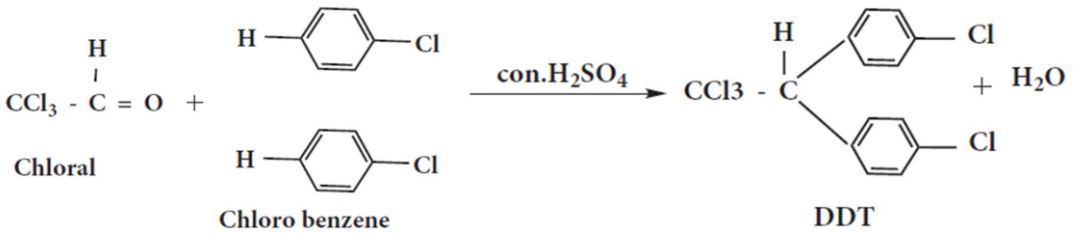
SN^1 reaction:

- Unimolecular nucleophilic substitution.
- Rate depends on concentration of alkyl halide.
- Follows first order kinetics.
- Occurs in two steps.
- Tertiary alkyl halides with aqueous KOH.

SN^2 reaction:

- Bimolecular nucleophilic substitution.
- Rate depends on concentration of alkyl halide and nucleophile.
- Follows second order kinetics.
- Occurs in one step.
- Primary alkyl halides with aqueous KOH.

2. How is DDT prepared?



3. Write the uses of DDT?

- Used to control malaria causing insect.
- Used to control agricultural pests.
- Used to kill mosquitoes.

4. How will you prepare n-propyl iodide from n-propyl bromide (or) What is Finkelstein reaction?

- $CH_3CH_2CH_2Br + NaI \xrightarrow{Acetone} CH_3CH_2CH_2I + NaBr$

5. What is Swarts reaction?

- $CH_3CH_2Br + AgF \xrightarrow{\Delta} CH_3CH_2F + AgBr$

6. Why chlorination of methane is not possible in dark?

- Free radical mechanism.
- Occurs only in light or heat.

7. How does chlorobenzene react with sodium in the presence of ether? What is the name of the reaction?

- Biphenyl is formed.
- Fittig reaction.
- $2\ C_6H_5Cl + 2Na \xrightarrow{Ether} C_6H_5C_6H_5 + 2NaCl$

8. Give reasons for polarity of C-X bond in haloalkane?

- Halogen is more electronegative than carbon.

9. Why is it necessary to avoid even traces of moisture during the use of Grignard reagent?

- Grignard reagents are very reactive.
- They react with moisture to form alkanes.

10. What happens when acetyl chloride is treated with the excess of CH_3MgI ?

- Tertiary butyl alcohol is formed.
- $CH_3COCl + CH_3MgI \xrightarrow{H_2O/H^+} CH_3COCH_3 \xrightarrow{CH_3MgI} (CH_3)_3COH$

11. What are Freons? Discuss their uses.

- The chloro fluoro derivatives of methane and ethane are called freons.

Uses:

- Used as refrigerants in refrigerators and air conditioners.
- Used as propellant for aerosols and foams.

UNIT 15. ENVIRONMENTAL CHEMISTRY

2 & 3 Mark questions

1. What is Acid Rain.

- The oxides of sulphur and nitrogen is absorbed by the water in the clouds and converted into sulphuric acid and nitric acid. This is called as Acid rain.
- The pH of rain water is 5.6

2. Define Smog.

- Smog is a combination of smoke and fog.

3. Which is considered to be earth's protective umbrella? Why?

- Ozone layer.
- It protects Earth from harmful UV radiations.

4. What are particulate pollutants?

- Particulate pollutants are small particles suspended in air.
- Ex. smoke, dust, mist.

5. What are degradable and nondegradable pollutants? Give examples.

Degradable pollutants:

- Pollutants that are easily decomposed by natural biological process.
- Ex. Animal and Plant wastes

Non degradable pollutants:

- Pollutants that are not easily decomposed by natural biological process.
- Ex. DDT, Plastics.

6. What is green chemistry?

- Green chemistry is a chemical philosophy encouraging the design of products and process that reduce the generation of hazardous substances.

7. What is greenhouse effect?

- The heating up of the earth's surface due to infrared radiations reflected by earth's surface by CO_2 in the atmosphere.

8. How is Greenhouse effect responsible for global warming?

- The heating up of earth's surface due to greenhouse effect is called global warming.

9. How does classical smog differ from photochemical smog?

Classical smoke:

- Occurs in cool humid climate.
- Occurs in the morning.
- Becomes worse when the sun rises.

Photochemical smog:

- Occurs in warm, dry and sunny climate.
- Forms when the sun shines.
- Becomes worse in the afternoon.

10. Differentiate BOD and COD?

BOD	COD
The total amount of oxygen in milligrams consumed by microorganisms by decomposing the waste in 1 liter of water at 20°C for a period of 5 days	The amount of oxygen required by the organic matter in a sample of water by $K_2Cr_2O_7$ in acid medium for 2 hours.
Expressed in ppm	Expressed in mg/l

11. Differentiate viable and non-viable particulate pollutants.

Viable particulate	Nonviable-particulate
The small size living organisms such as dispersed in air. Ex. Bacteria, fungi.	The small solid particles and liquid droplets suspended in air. Ex. Smoke, dust.
They contain living organisms	They contain non living organisms.

12. Mention the standards prescribed by BIS for quality of drinking water.

S.No	Characteristics	Desirable limit
1.	pH	6.5 to 8.5
2.	TDS	500 ppm
3.	Total Hardness	300 ppm