LEARNING ENHANCEMENT PROGRAM

STUDY MATERIAL 2025-26

CLASS 11TH

(CHEMISTRY)

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Chapter 1 - Some Basic Concepts of Chemistry

Chemistry is the branch of science that studies the composition, properties, structure of matter and changes which the matter undergoes under different conditions and laws which govern these changes.

❖ Anything that occupies space and has mass is matter.

❖ The term precision refers for the closeness of the set of values obtained from identical measurements of a quantity.

❖ The term accuracy refers to be closeness of a single measurement to its true value.

❖ All numbers, small or large are expressed as a number between 1.000 and 9.999 multiplied or divided by 10, an appropriate number of times.

❖ 1426.2 = 1.4262 × 10^3 in scientific rotation, a number is generally expressed in the form N × 10^n. Where N is a number (called digit term) between 1.000 and 9.999 and *n* is a number called an exponent.

❖ The significant figures in a number are all the certain digits plus one doubtful digits.

❖ In addition and subtraction, the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places.

❖ In multiplication and division the final results should be reported as having the same number of significant digits as the number with least number of significant digits.

❖ One twelth $\left(\frac{1}{12}\right)$ of the mass of an atom of carbon (C – 12) is atomic mass unit. It is equal to 1.66 × 10^{-27} kg.

❖ The average relative mass of an atom of element as compared with mass of a carbon atom taken as 12 a.m.u. is atomic mass.

❖ The average relative mass of a molecules of the substance as compared with mass of a carbon atom taken as 12 a.m.u. is molecular mass.

❖ Atomic mass expressed in grams is gram atomic mass and molecular mass expressed in grams is gram molecular mass.

Chapter 2 :- Structure of Atom

❖ According to Dalton atomic theory matter is the smallest or ultimate particle of an atoms.

❖ Cathode rays consist of negatively charged particle called electrons.

❖ An electron has –1 unit charge (1.6 × 10^{-19} C) and negligible mass (9.1 × 10^{-31} kg).

❖ An electron is regarded as an universal particle.

❖ Anode rays also called canal rays consist of positively charged known as proton.

❖ Anode rays are not emitted from the anode but from a space between anode and cathode.

❖ A proton is not a universal particle like electron.

❖ A proton has +1 unit charge (1.6 × 10^{-19} C) which has same magnitude as an electron and 1 *u* mass (1.67 × 10^{-27} kg).

❖ The phenomenon of spontaneous emission of certain highly active radiation by radioactive substance is called radioactivity.

❖ Thomson's model of atom regards an atom as a positively charged sphere in which negatively charged electrons are embeded. It is called Plum Pudding model of an atom.

❖ Alpha particle are helium nuclei $^{4}_{2}He$. They have 4 unit mass and +2 unit charge.

❖ Rutherford's model of atom is based on α-ray scattering experiment. According to it, an atom consist of a very small positively charged nucleus containing protons around which negatively charged electrons are distributed in a space known as extra-nuclear portion.

❖ Alpha particle for scattering experiment are obtained from radioactive element radium.

❖ Electrons are also known as planetary electrons.

❖ Neutrons were discovered by Chadwick. There are neutral particles present in the nucleus along with protons.

❖ Electromagnetic spectrum is the arrangement of different electromagnetic radiations in order on increasing wavelength or decreasing frequency.

Chapter 3: Classification of elements and periodicity in properties

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

❖ "The physical and chemical properties of elements are periodic functions of their atomic weight."

❖ 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 $\rightarrow$ Periods

Vertical columns $\rightarrow$ Group

❖ 1st period – 2 elements

2nd and 3rd period – 8 elements

4th and 5th period – 18 elements

6th period – 32 elements

7th period – Incomplete

❖ 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 '*f*'subshell [p^1 to p^6].

Group 18 – Noble gases.

❖ (*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 sub penultimate shell (n – 2).

(*b*) '*f*' block elements are placed in between '*d*' block elements.

'*f*' block elements in 2 rows [4 *f* lanthanoids, 5 *f* actinoids]

(*c*) Helium is placed ns^2. But it has Noble gas configuration.

❖ General electronic configuration '*s*' block ns^1, ns^2 [Group I]

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

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

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

❖ General Trends :

Atomic Radius :

(*a*) Left to right decreases due to effect of successive increasing nuclear change without addition of a new shell.

(*b*) From top to bottom atomic radius increases due to successive addition of shell.

(*c*) Noble gases larger adius than group 17 due to complete filling of electron in outer shell electron-electron repulsion mildy increases.

CHAPTER 4: CHEMICAL BONDING AND MOLECULAR STRUCTURE

❖ The interaction between two atoms which holds them together within a molecule or ions in known as chemical bond.

❖ The elements with one, two, three, four, five, six or seven electrons is outer shell, use these electrons to complete octet. The electrons which take part in two or more atoms to complete octet is known as electro valency.

❖ Lewis symbols or electron dot symbols involve the presentation of valence electrons (outer electrons) in an atom $\dot{\mathrm{Na}},\ \mathrm{Mg}\!:,\ \underset{\cdot}{\ddot{\mathrm{Al}}},\ \underset{\cdot\cdot}{\ddot{\mathrm{Si}}}$ etc.

❖ 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) [$2e^-$ in H, Li] is known as ionic bond.

❖ 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.

❖ 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 H_2O [Hydration].

❖ Lattice enthalpy : The energy released when one more of ionic compound is formed from its ions in their gaseous state. Lattice energy is directly proportional to charge of ion *nd* inversely proportions to size of ions *i.e.*, more is change density more is lattice energy.

Fajan's Rule : Polarizability and polorizing power. The power of anion to distart the other ion is polarising power and the tendency of an ion to get distorted is known as polarizability. Factors affect ion polarizing power and polarizability.

(*a*) High charge and small size of C^+.

(*b*) High charge and large size of A^-.

❖ 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 covalence bond and no. of electrons involved is covalency.

❖ Formal charge : Electron of an atom in a molecule/ion

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

CHAPTER: 5 (CHEMICAL THERMODYNAMICS)

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 : $\Delta T = 0$ (T = constant)

(*ii*) Adiabatic process : $dq = 0$ (Heat = constant)

(*iii*) Isobaric process : $\Delta P = 0$ (P = constant)

(*iv*) Isochoric process : $\Delta V = 0$ (V = constant)

(*v*) Cyclic process : $\Delta U = 0$ (U = constant)

(*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.

CHAPTER 6: EQUILIBRIUM

The state of which the measurable properties of the system do not undergo any further notice-able change under given set of condition is said to be a state of equilibrium.

❖ At equilibrium rate of forward reaction becomes equal to rate of backward reaction. Also at equilibrium, ΔG becomes equal to zero so that $\Delta H = T\Delta S$.

❖ Chemical reaction reach a state of dynamic equilibrium in which the rate of forward and backward reaction are equal and there is no net change in composition.

❖ Homogeneous equilibrium in which reactants and products are in the same phase.

❖ Heterogeneous equilibrium in which various substances involved in the reaction have different phases.

❖ The Main Characteristics of The Equilibrium : It can be attained if the system is closed; It is dynamic in nature; It can be approached from either direction; a catalyst does not alter the equilibrium point.

❖ Law of Mass Action : The rate of a chemical reaction at any particular temperature is proportional to the product of the molar concentration of reactants with each concentration term raised to the power equal to the number of molecules of the respective reactants.

❖ Equilibrium Constant : It is the ratio between the products of molar concentration of the reactants, which each concentration term raised to a power equal to its stoichiometric coefficients in the balanced chemical equations at a constant temperature.

❖ The units of equilibrium constant are units based on molarity or pressure, unless the sum of the exponents in the numerator is equal to the sum of the exponents in the denominator.

❖ Le-Chatelier's Principle : If an equilibrium is subjected to any kind of stress equilibrium shifts in such a way so as to undo the effects of stress imposed.

❖ The equilibrium established between the unionized molecules and the ions in solution of weak electrolytes is or degree of ionization or degree of dissociations.

CHAPTER 7: REDOX REACTION

❖1. Fundamental Concepts

- Oxidation: Loss of electrons, increase in oxidation number (O.N.), addition of O/electronegative element, removal of H/electropositive element.
- Reduction: Gain of electrons, decrease in O.N., addition of H/electropositive element, removal of O/electronegative element.
- Redox: Simultaneous occurrence of oxidation and reduction.
- Oxidizing Agent (Oxidant): Substance that gets reduced (accepts electrons).
- Reducing Agent (Reductant): Substance that gets oxidized (donates electrons).

❖2. Oxidation Number (O.N.)

- A number representing the apparent charge on an atom in a compound, following set rules.
- Key Rules: F is always -1; O is usually -2 (except peroxides, superoxides); H is +1 (except metal hydrides); Elements in free state are 0; Total O.N. in neutral molecule is 0; Total O.N. in ion equals ion's charge.

❖3. Types of Redox Reactions

- Combination: Two or more reactants form one product (e.g.,

$$2Na + Cl_2 \rightarrow 2NaCl$$

4. Balancing Redox Reactions

- Oxidation Number Method: Adjust coefficients to balance atoms and the *change* in total O.N..

- Ion-Electron (Half-Reaction) Method: Separate into oxidation and reduction half-reactions, balance atoms and charge with
    - Acidic Medium: Add H^+ and H_2O FOR O BALANCE
    - Basic Medium: Add OH^- For O and H_2O for H balance.

CHAPTER : 8 Organic Chemistry

Functional Group : Determines the characteristic properties of a compound.

❖ Functional group may be carbon carbon multiple bond or carbon bonded to other atoms such as N, O, S or P.

❖ Homologous Series : is a series of similarly constituted compounds in which the members possess the same functional group and have similar chemical characteristics.

❖ The two consecutive members of a homologous series differ in their molecular formula by – CH_2 group.

❖ Isomers : are the compound which have the same molecular formula but differ in their physical and chemical properties.

❖ Structural Isomers : differ from one another in the arrangement of atoms or group of atoms within the molecules.

❖ Inductive Effect : is the process of electron displacement along chain of carbon atoms due to presence of a polar covalent bond at one end of the chain.

❖ Electromeric Effect : is the phenomenon of movement of electron from one atom to another in multi bonded atoms at the demand of attacking reagent.

❖ If a molecule can be represented by two or more structure none of which is capable of describing all the known properties of the compound, then the actual structure is intermediate or resonance hybrid of these structure. Such phenomenon is called resonance.

CHAPTER:9 HYDROCARBON

Classification of Hydrocarbons

- Saturated (Alkanes): Single C-C bonds (e.g., Methane, Ethane).
- Unsaturated (Alkenes/Alkynes): Double or triple C-C bonds (e.g., Ethene, Ethyne).
- Aromatic (Arenes): Cyclic, stable with alternating double bonds (e.g., Benzene).

2. Alkanes (Saturated)

- General Formula: CnH_{2n+2}
- Isomerism: Chain, position, conformational (eclipsed/staggered).

- Preparation: Hydrogenation, Wurtz Reaction, Decarboxylation.
- Properties: Relatively unreactive, undergo free radical substitution (halogenation), combustion.

3. Alkenes (Unsaturated)

- General Formula: C_nH_{2n} sp^2 ,hybridization, trigonal planar geometry.
- Isomerism: Geometrical (cis-trans).
- Reactions: Electrophilic addition (Markownikoff's Rule, Anti-Markownikoff), hydration, polymerization.

4. Alkynes (Unsaturated)

- General Formula: CnH_{2n-2}

.sphybridization, linear geometry.

- Reactions: Electrophilic addition (less reactive than alkenes), hydration, polymerization.

5. Aromatic Hydrocarbons (Arenes)

- Benzene C_6H_6

: Cyclic, planar, resonance, aromaticity (Hückel's Rule).

- Reactions: Electrophilic Substitution (Nitration, Halogenation, Sulphonation, Friedel-Crafts Alkylation/Acylation).
- Directive Influence: Ortho-, Para-, Meta-directors.

6. Important Concepts & Uses

- Isomerism: Crucial for distinguishing isomers (e.g., N-pentane vs. Neopentane M.P./B.P.).
- Conformational Isomerism: Eclipsed vs. Staggered (Ethane).
- Uses: Fuels (LPG, CNG, Petrol), solvents, polymers (Polythene).
- Toxicity/Environmental Impact: Mentioned in context of fuels.

Chapter 1 - Some Basic Concepts of Chemistry

MCQs (1 Mark Each)

1. SI unit of amount of substance is
   A. gram B. mole C. kg D. litre
2. 1 mole contains
   A. 6.022×10^{22} particles B. 6.022×10^{23} particles
   C. 3.011×10^{23} particles D. 1.66×10^{-24} particles
3. Law of conservation of mass was given by
   A. Dalton B. Lavoisier C. Gay-Lussac D. Avogadro
4. Empirical formula shows
   A. Actual atoms B. Simplest ratio C. Molecular mass D. Structural formula
5. 22.4 L of any gas at STP contains
   A. 1 mole B. 2 moles C. 0.5 mole D. 44.8 moles
6. Molar mass of CO_2 is
   A. 28 g B. 32 g C. 44 g D. 16 g
7. Limiting reagent determines
   A. Rate B. Yield C. Volume D. Temperature
8. Percentage purity depends on
   A. Impurities B. Mass C. Temperature D. Pressure
9. Significant figures in 0.00450
   A. 2 B. 3 C. 4 D. 5
10. Law of definite proportions was proposed by
    A. Dalton B. Proust C. Lavoisier D. Avogadro

Answer Key (Ch-1)

1-B, 2-B, 3-B, 4-B, 5-A, 6-C, 7-B, 8-A, 9-B, 10-B

2 Marks Question Answers

1. Define law of multiple proportions with example.
2. Calculate the molecular mass of $C_{12}H_{22}O_{11}$
3. Calculate the no. of atoms present in 11.5 litres of H_2 at N.T.P.
4. Calculate the no. of moles of 5.68 gm. of iron.
5. What is the effect of temp. on molality and molarity?
6. An atom of an element is 10.1 times heavier than the mass of a carbon atom. What is its mass in a.m.u.?
7. Explain with example, limiting reagent.
8. Differentiate between molarity and molality.

9. 1.82 g. of glucose (molar mass-180) is dissolved in 25g of water. Calculate (a) the molality (b) mole fraction of glucose and water.
10. The molecular mass of an organic compound is 90 and its %age composition is C- 26.6%; O=71.1% and H=2.2%. Determine the molecular formula of the compound.
11. How chemical equations are made more informative?
12. How Avogadro's hypothesis used to deduce atomicity of elementary gases?
13. Verify law of Reciprocal proportions or law of equivalent proportions, with example.
14. Define formula mass and how does it differs from molecular mass?
15. Discuss Dalton's Atomic theory and its limitations?
16. Discuss Modern Atomic theory. Why it is better than Dalton's Atomic theory?
17. Commercially available sulphuric acid contains 91% acid by mass and has a density of 1.83g mL^{-1} (i) Calculate the molarity of the solution (ii) volume of concentrated acid required to prepare 3.5L of 0.50 M H_2SO_4

3 Marks Question Answers

1. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% of chlorine. Its molar mass is 98.96g. What are its empirical and molecular formulas?
2. How much copper can be obtained from 110gm of $CuSO_4$?
3. What is Gay Lussac's law? Explain with two examples.
4. What are empirical and molecular formulae? How are they related to each other?
5. Differentiate between normality and molarity?
6. Why molality is preferred over molarity in expressing the concentration of a solution?
7. Explain with the help of an example law of conservation of mass and energy and also the law of constant proportions.
8. What is the law of conservation of mass? Give one example.
9. What are the main postulates of Dalton's atomic theory?
10. Calculate the number of Cu atoms in 0.635 g of Cu.
11. 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?
12. Calculate the number of atoms in each of the following a) 52 moles of Ar b) 52 u of He and c) 52 g of He

Chapter 2 :- Structure of Atom

Multiple Choice Question each (1 Marks):

1. Discovered electron
   A. Rutherford B. Bohr C. J.J. Thomson D. Chadwick
2. Charge of electron is
   A. +1 B. −1 C. 0 D. +2
3. Maximum electrons in M-shell
   A. 8 B. 18 C. 32 D. 50
4. Dual nature of matter given by
   A. Einstein B. Bohr C. de Broglie D. Planck
5. Neutrons were discovered by
   A. Chadwick B. Bohr C. Thomson D. Goldstein
6. Which has highest energy?
   A. 1s B. 2s C. 2p D. 3s
7. Shape of s-orbital
   A. Dumbbell B. Spherical C. Linear D. Square planar
8. Maximum electrons in p-subshell
   A. 2 B. 6 C. 10 D. 14
9. Atomic number equals
   A. Neutrons B. Protons
   C. Mass number D. Protons + neutrons
10. Pauli exclusion principle related to
   A. Energy B. Spin C. Orbit D. Charge

Answer Key (Ch-2)

1-C, 2-B, 3-C, 4-C, 5-A, 6-D, 7-B, 8-B, 9-B, 10-B

Short Answer type Question each (2 Marks)

1. From the following nuclei select the isotopes and isobars.

   $^{238}_{92}U$, $^{234}_{90}Th$, $^{234}_{92}U$, $^{234}_{91}Pa$, $^{238}_{93}Np$

2. What is Zeeman effect and Stark effect?
3. Write electronic configurations, of Cr, Cu, Zn?
4. Define Aufbau's Principle. Which of the following orbitals are possible. 1 s, 1 p, 2 s, 3 d, 3 f
5. Explain Hund's rule of maximum multiplicity by taking an example of phosphorous.
6. Why are Bohr's orbits called Stationary States?
7. What is the difference between atomic mass and mass number?
8. Explain why the uncertainty principle is significant only for the microscopic particles and not for the macroscopic particles?
9. Why half-filled and fully filled orbitals are extra stable?

10. Why electronic configuration of 'Cr' is $3d^5\ 4s^1$ and not $3d^4\ 4s^2$ and 'Cu' is $3d^{10}\ 4s^1$ and not $3d^9\ 4s^2$?

11. Give differences between orbit and orbital.

12. What is photoelectric effect? What is the effect of frequency and intensity on photoelectric effect?

13. Why large no. of lines appear in the spectrum of hydrogen although it contains only one electron?

14. Derive de Broglie relationship and give its significance.

15. Give important postulates of Bohr's model of an atom.

16. Discuss Planck's Quantum theory of Radiation.

17. Using the s, p, d, f, notations describe the following quantum no.

18. (a) n =1, l =0 (c) n = 4; l =3 (d) n = 4; l = 2

19. (b) n = 3, l =2 (d) n = 5; l = 4 (e) n = 6; l = 4

Short Answer type Question each (3 Marks)

1. The electronic configuration of the valence shell of Cu is $3d^{10}\ 4s^1$ and not $3d^9\ 4s^2$. How is this configuration explained?
2. What is the experimental evidence in support of the idea that electronic energies in an atom are quantized?
3. Out of electrons and protons which one will have a higher velocity to produce matter waves of the same wavelength? Explain it.
4. What is the difference between the terms orbit and orbital? Discuss important facts about photoelectric effect.
5. Discuss black body radiation. Also explain its reason.
6. What are emission and absorption spectra? Why dark lines appear in the absorption spectra?
7. What is the frequency and wavelength of a photon emitted during a transition from n=5 state to n=2 state in the hydrogen atom.
8. Discuss drawbacks of Rutherford's Model.
9. Explain Heisenberg's uncertainty Principle.
10. What do you understand by an atomic orbital? Briefly describe the shapes of s, p & 'd' orbitals?
11. State and explain Aufbau's principle, Pauli's exclusion principle.

z

Chapter 3: Classification of elements and periodicity in properties

1. What are magic numbers?
2. Give Modern periodic law.
3. What are Dobereiner's triads?
4. Give general electronic configuration of 'd'-block and 'f'-block elements.
5. What are the defects of long form of the periodic table?
6. What is the cause of periodicity?
7. What are successive Ionization enthalpies?
8. Why Ionization enthalpy of 'Be' is more than 'B' and of 'N' is more than 'O' explain?
9. Why electron gain enthalpies of Noble gases are positive while those of 'Mg' and 'P' are almost zero?
10. Why electron gain enthalpy of flourine is less negative than that of chlorine?
11. What are iso electronic species? How are their sizes vary in iso electronic series?
12. Which of the following will have the largest and smallest size and why? Cl, Cl^{-1}, Al, Al^{3+}
13. Why d- and f-block elements are less electropositive than group 1 and 2 elements?
14. What is diagonal relationship? Explain it with the help of 'Be' and 'Al'.
15. What is ionisation enthalpy? On what factors it depends?
16. What is electron gain enthalpy? On what factors it depends. How it varies in a group and in a period?
17. How will you justify presence of 18 elements in 5^{th} period and presence of 32 elements in 6^{th} period?

Comprehension-:

1. There are different periodic trends with respect to atomic and ionic radii. Size of an atom non metallic element is to measure the distance between two atoms when they are bound together by a single bond in covalent molecule. Atomic radius refer to both covalent or metallic radius depending on whether the element is non metal or metal. The removal of an electron from an atom resulting in the formation of cation, where as gain of an electron leads to an anion. The ionic radii can be estimated by measuring the distance between cations and anions in ionic crystals.

1. Which of the following is the correct order of the size of given species ?

(A) $I > I^+ > I^-$ (B) $I^- > I > I^+$ (C) $I^+ > I > I^-$ (D) $I^- > I^+ > I$

2. Which of the following has smallest size?
(A) N (B) B (C) C (D) F
3. Which of the following order of ionic radii is correctly is represented ?
(A) $O^{2-} > F^- > Na^+ > Mg^{2+}$ (B) $Mg^{2+} > Na^+ F^- > O^{2-}$
(C) $F^- > O^{2-} > Mg^{2+} > Na^+$ (D) $Na^+ > Mg^{2+} > O^{2-} > F^-$
4. The atomic radii of Be and Al respectively are

(A) 143 pm and 111 pm (B) 111 pm and 111 pm
(C) 111 pm and 143 pm (D) 143 pm and 143 pm

2. In the modern periodic table, element are arranged in order of increasing atomic number which is related to the electronic configuration. Depending upon the type of orbitals receiving the last electron, the element in which periodic table have been divided into four blocks, viz***, s, p, d and f.*** the modern periodic table consists of 7 periods and 18 groups. Each period begins with the filling of a new energy shell. In accordance with the Aufbau principal, the 7 periods (1to 7) have 2, 8, 8, 18, 18, 32 and 32 elements respectively. The seventh period is still incomplete. To avoid the periodic table being too long, the two series of ***f***-block elements, called lanthanoids and actinoids are placed at the bottom of the main body of the periodic table .

1. The element with atomic number 57 belong to
   (a) S-block
   (b) P-block
   (c) d-block
   (d) f-block
2. The last element of *p* block in 6th period is represented by the outer most electronic configuration
   (a) $7s^2 7p^6$
   (b) $5f^{14} 6d^{10} 7s^2 7p^0$
   (c) $4f^{14} 5d^{10} 6s^{10} 6p^{6}$
   (d) $4f^{14} 5d^{10} 6s^2 6p^4$
3. Which of the elements whose atomic number are given below, cannot be accommodated in the present set up of the long form of the periodic table ?
   (a) 107
   (b) 118
   (c) 126
   (d) 102
4. The elements with atomic number 35, 53 and 85 are all ___________
   (a) Noble gases
   (b) Halogens
   (c) Heavy metals
   (d) Light metals
5. Which of the following is correct order of increasing tendency to gain electrons;
   (a) A<C<B<D
   (b) A<B<C<D
   (c) D<B<C<A
   (d) D<A<B<C

CHAPTER 4: CHEMICAL BONDING AND MOLECULAR STRUCTURE

MCQs :

1. Ionic bond formed by
   A. Sharing B. Transfer C. Overlap D. Hybridisation

2. Bond angle of CH_4
A. 90° B. 120° C. 109.5° D. 180°
3. Hybridisation in C_2H_4
A. sp B. sp^2 C. sp^3 D. dsp^2
4. Strongest bond
A. Ionic B. Hydrogen C. Covalent D. Metallic
5. Shape of NH_3
A. Linear B. Trigonal planar C. Pyramidal D. Tetrahedral
6. VSEPR theory explains
A. Bond length B. Bond angle C. Energy D. Charge
7. Hydrogen bond strongest in
A. NH_3 B. H_2O C. HF D. HCl
8. Octet rule violated by
A. C B. O C. BF_3 D. N
9. σ bond formed by
A. Side overlap B. End overlap C. p-p overlap D. d-d overlap
10. Polar bond formed due to difference in
A. Mass B. SizeC. Electronegativity D. Valency

Answer Key (Ch-4)

1-B, 2-C, 3-B, 4-A, 5-C, 6-B, 7-C, 8-C, 9-B, 10-C

TWO OR THREE MARKS QUESTIONS :

1. Why do atoms combine?
2. What is the significance of Lewis Symbols?
3. Give structure of BrF_5
4. Why H_2O is liquid and H_2S is a gas?
5. Why NH_3 is liquid and PH_3 is a gas?
6. Boiling point of p-nitrophenol is more than O-nitrophenol why?
7. How is paramagnetic character of a compound is related to the no. of unpaired electrons?
8. Describe a co-ordinate bond with an example. How does it differs from a covalent bond?
9. How is MgF_2 and Al_2O_3 formed?
10. What is an Octet rule? What are its limitations?
11. Which out of NH_3 and NF_3 has higher dipole moment and why?
12. Draw molecular orbital diagram for N_2^+ molecule.
13. HCl is a covalent compound but it ionises in the solution?
14. The molecule of CO_2 is linear whereas that of $SnCl_2$ is angular why?
15. Give molecular orbital energy level diagram of CO. Write its electronic configuration, magnetic behaviour and bond order.

16. How is ionic bond formed? On what factors it depends?

17. Calculate the lattice enthalpy of KCl from the following data by Born-Haber's Cycle. Enthalpy of sublimation of K=89 KJ mol^{-1} ,Enthalpy of dissociation of Cl_2 = 244 KJ mol^{-1} Ionization enthalpy of potassium = 425 KJ mol^{-1} Electron gain enthalpy of chlorine = - 355 KJ mol^{-1} Enthalpy of formation of KCl = -438 KJ mol^{-1}

FIVE MARKS QUESTIONS

1. Using a molecular orbital diagram , predict the bond order , stability and magnetic character of O_2^-,O_2^+ and O_2^{2-} ions . Also write their electronic configuration .
2. Draw the molecular orbital energy diagram for oxygen molecule (O_2) and show that :
   (i) it has a double bond
   (ii)it has paramagnetic character.
3. Use LCAO method for the formation of molecular orbitals in case of homonuclear diatomic hydrogen molecule.
4. Using a molecular orbital diagram , predict the bond order , stability and magnetic character of N_2^-,N_2^+ and N_2^{2-} ions . Also write their electronic configuration .

CHAPTER: 5 (CHEMICAL THERMODYNAMICS)

MCQS

1. A thermodynamic state function is a quantity :
   (A) Use to determine heat changes.
   (B) Whose value is independent of path
   (C) Use to determine pressure volume work
   (D) Whose values depends in temperature only.

2. The work done in case of isothermal free expansion is:
   (A) Maximum (B) Minimum (C) Zero (D) Positive
3. The enthalpies of all elements in their standard state are equal to:
   (A) Unity (B) Zero (C) <0 (D) Different for each element
4. Thermodynamics is not related to:
   (A).Energy changes evolved in a chemical reaction
   (B) The extent to which a chemical reaction proceeds
   (C) The rate at which a reaction proceeds
   (D).The feasibility of a chemical reaction
5. Which of the following is extensive property ?
   (A) Molar heat capacity (B) Temperature (C) Enthalpy (D) All of these
6. The enthalpies of elements in their standard states are taken as Zero. The Enthalpy of formation of a compound
   (A) Is always negative (B) Is always positive
   (C) May be positive or negative (D) Is never negative
7. Enthalpy of sublimation of a substance is equal to:
   (A) Enthalpy of fusion+ Enthalpy of vapourisation (B) Enthalpy of fusion
   (C) Enthalpy of vapourisation (D) Twice the Enthalpy of vapourisation
8. If the volume of gas is reduced to half of its original volume then the specific heat will be _____.
   (A) reduce to half (B) be doubled (C) remain constant (D) increase four times
9. A well stoppered thermos flask contain some ice cubes. This is an example of
   (A) Closed system (B) Open system (C) Isolated system (D) Non thermodynamics system
10. The Enthalpy of vaporisation of a substance is 8400 J mol-1 and its boiling point is -173°C. The entropy change for vaporisation is :
    (A) 84 mol-1 k-1 (B) 21 mol-1 k-1 (C) 49 mol-1 k-1 (D) 12 mol-1 k-1

11. A system absorbs 10 KJ of heat at constant volume and its temperature rises from 270 to 370 C . The volume of delta U is
(A) 100 KJ (B) 10 KJ (C) 0 KJ (D) 1 KJ

2 MARKS QUESTIONS

1. Name the different types of system.
2. What will happen to internal Energy if work is done by the system ?
3. State second law of thermodynamics?
4. Define enthalpy
5. Give the mathematical expression of first law of thermodynamics.
6. Define Hess law. Why this law is called law of constant heat summation.
7. Under what conditions enthalpy change and internal energy change of system remains same?

3 MARKS QUESTIONS

1. Calculate enthalpy formation of ethane from the following data:
   $C + O_2 \rightarrow CO_2$ ($\Delta H = -393.5$KJ /mol)
   $H_2 + O_2 \rightarrow H_2O$($\Delta H$=-285.8kJ/mol)
   $C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$($\Delta H$=-1560KJ/mol)

2. Define the following terms:
   Internal energy ,entropy, adiabatic process , isochoric process
3. Absolute value of internal energy cannot be determined. Explain.
4. When Δ G is positive, the process is always non spontaneous. Explain.
5. Explain the meaning of driving force of a chemical reaction. How is Δ G related to Δ H and Δ S in a reaction?
6. How does T Δ S determine the spontaneity of a process?
7. How will you justify that both 'q' and 'w' are not state functions, yet (q+w) is a state function?
8. Δ H is negative for exothermic reaction and positive for endothermic reaction. Explain.
9. For a reaction both Δ H and Δ S are positive. Under what conditions will the reaction be spontaneous?
10. Determine Δ H° at 298 K for the reaction.
    C(graphite) + $2H_{2(g)} \rightarrow CH_{4(g)}$; ΔH^o = ? you are given
    C(graphite) + $O_{2(g)} \rightarrow CO_{2(g)}$ ΔH^o = - 393.5 KJ mol^{-1}
    $H_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$ $\Delta H^o r$ = -285.8 KJ mol^{-1}
    $CO_{2(g)} + 2H_2O_{(l)} \rightarrow CH_{4(g)} + 2O_2(g)$; ΔH^o = +890.3 KJ mol^{-1}
11. Predict the feasibility of a reaction when both Δ H and Δ S are negative.
12. For the reaction $A_{(g)} + 3B_{(g)} \rightarrow 2C_{(g)}$, the enthalpy change is -90.2 KJ mol^{-1} and Δ S is - 0.1584 KJ K^{-1} mol^{-1}. Predict whether the reaction is feasible or not at 298 K?
13. Enthalpy and entropy changes of a reaction are 49.57 KJ mol^{-1} and 123.2 J K^{-1} mol^{-1}. Calculate the free energy change of the reaction at 27°C.

CHAPTER 6: Equilibrium

MCQS

1. Chemical equilibrium is
   A. Static B. Dynamic C. Unidirectional D. Complete
2. Kc is equilibrium constant in terms of
   A. Pressure B. Mole fraction C. Concentration D. Volume
3. Value of K depends on
   A. Concentration B. Pressure C. Temperature D. Catalyst
4. Catalyst affects
   A. K B. ΔH C. Rate D. Equilibrium position
5. pH < 7 indicates
   A. Neutral B. Basic C. Acidic D. Salt
6. Weak acid has
   A. High Ka B. Low Ka C. Ka = 1 D. No Ka
7. Le Chatelier principle predicts effect of
   A. Catalyst B. Temperature C. Stress D. Time
8. Ionic product of water at 25°C
   A. 10^{-7} B. 10^{-14} C. 10^{-12} D. 10^{-6}
9. Buffer solution resists change in
   A. Temperature B. Pressure C. pH D. Volume
10. Strong base completely
    A. Ionizes B. Hydrolyses C. Reacts D. Neutralizes

Answer Key (Ch-7)

1-B, 2-C, 3-C, 4-C, 5-C, 6-B, 7-C, 8-B, 9-C, 10-A

Two or Three mark questions

1. What do you mean by common ion effect? Explain.
2. State and explain Le-chatelier's principle.
3. What will be pH of 0.001M NaOH solution?
4. The equilibrium constant expression for a gas reaction is
   $$K_c = [NH_3]^4 [O_2]^5] / [NO]^4 [H_2O]^6$$
   Write the balanced chemical equation corresponding to this expression.
5. The solubility of A_2X_3 is Y mol dm^{-3}. Calculate its solubility product
6. Give limitations of Arrhenius concept of Acids and bases.
7. Give advantages of Bronsted-Lowry concept over Arrhenius concept.
8. (i) What will be the conjugate bases for the following Bronsted acids? HF, H_2SO_4, HCO_3, H_3PO_4

   (ii) Why PO_4^{3-} ion is not amphoteric?
9. What is a buffer solution? Ammonium acetate is a buffer where as sodium chloride is not. Why?
10. What are acidic buffers? Explain with the help of an example.
11. What are basic buffers? Explain with the help of an example.

CHAPTER 7: REDOX REACTION

MULTIPLE CHOICE QUESTIONS:

1. Oxidation involves
   A. Gain of electrons B. Loss of electrons
   C. Gain of neutrons D. Loss of protons
2. Reduction involves
   A. Gain of electrons B. Loss of electrons
   C. Gain of oxygen D. Loss of hydrogen
3. Oxidizing agent
   A. Gets oxidized B. Gets reduced
   C. Loses electrons D. Gains electrons and oxidizes others
4. Redox reaction involves
   A. Only oxidation B. Only reduction C. Both D. Neither
5. Oxidation number of O in H_2O_2
   A. −2 B. −1 C. 0 D. +1
6. Oxidation number of Na in NaCl
   A. 0 B. −1 C. +1 D. +2
7. Which is redox reaction?
   A. Neutralization B. Precipitation C. Combination D. Double displacement
8. $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ In this reaction Zn is
   A. Reduced B. Oxidized C. Neutral D. Catalyst
9. Disproportionation involves
   A. Only oxidation B. Only reduction
   C. Same element oxidized & reduced D. Two elements
10. Rusting of iron is
    A. Reduction B. Oxidation C. Neutralization D. Displacement

Answer Key (Ch-8)

1-B, 2-A, 3-D, 4-C, 5-B, 6-C, 7-C, 8-B, 9-C, 10-B

TRUE / FALSE:

1. When a substance is oxidized, it loses electrons.
2. Oxidation is loss of oxygen atom or gain of hydrogen atom.
3. The reducing agent is oxidized.
4. The oxidizing agent is reduced.
5. Reduction can also be defined as the gain of oxygen atoms or the loss of hydrogen atoms.

Two mark question:

1. Define oxidation reaction.
2. Define oxidation in terms of electron transfer.
3. Define reduction reaction.
4. What is the oxidation number of manganese in KMnO4?
5. Define an oxidizing agent. Name the best reducing agent.
6. (i) Why are redox reactions called electron transfer reaction?

   (ii) Can the same element have different oxidation numbers in different compounds? Justify.
7. (i) What happens when a zinc rod is dipped in a copper sulphate solution?

   (ii) What are combination redox reactions and decomposition redox reactions? Give examples.
8. H_2S acts as a reducing agent while SO_2 acts as an oxidising as well as reducing agent. Explain.
9. Give important features of Half-cell reactions.
10. HNO_3 acts as an oxidising agent while HNO_2 can act both as a reducing agent as well as oxidising agent explain.
11. Give differences between oxidation no. and valency.
12. Are all decomposition reactions redox reactions? Comment.

Three mark questions:

1. Explain electrochemical cell and electrolytic cell with the help of diagrams.
2. Oxidation cannot occur without reduction. Justify.
3. Explain standard hydrogen electrode (SHE) with diagram.
4. Explain oxidation and reduction according to electronic concept. Give two examples.
5. Discuss redox reactions on the basis of oxidation number.
6. Balance following equations by oxidation no. method.
   a. $SnO_2 + C \rightarrow Sn + CO$
   b. $Zn + NO_3^- + H^+ \rightarrow Zn^{2+} + N_2O + H_2O$
   c. $NH_3 + O_2 \rightarrow NO + H_2O$
7. Balance following equations by Ion-Electron method.
   a. $Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightarrow Cr^{3+} + Fe^{3+} + H_2O$
   β. $NO_3^- + Zn \rightarrow Zn^{2+} + NH_4^+$
8. Give differences between Electrochemical cell and Electrolytic cell.
9. What are disproportionation redox reaction? Give example.
10. Give limitations of concept of oxidation number.
11. Give advantage of electron density concept over oxidation no. concept.
12. Discuss the role of redox titrations in volumetric titrations.
13. Chlorine, bromine and iodine disproportionate in alkaline medium but fluorine does

not. Why?

14. Give an important application of non-metal displacement redox reactions in qualitative mixture analysis.

CHAPTER : 8 Organic Chemistry

1. What is the IUPAC name of the compound CH_3-CH_2-CH_2-CH_3?
   a) Propane b) Butane c) Pentane d) Hexane
2. The functional group -COOH is known as:
   a) Alcohol b) Aldehyde c) Carboxyl d) Ketone
3. What is the hybridization of the carbon atom in methane (CH_4)?
   a) sp b) sp^2 c) sp^3 d) sp^3d
4. Which of the following compounds is an example of an alkene?
   a) Ethane b) Ethene c) Ethyne d) Methane
5. Which of the following reactions is not possible in organic chemistry?
   a) Addition reaction b) Substitution reaction
   c) Combustion reaction d) Nuclear reaction
6. Which type of isomerism is exhibited by 1-butene and 2-butene?
   a) Structural isomerism b) Geometrical isomerism
   c) Optical isomerism d) Functional isomerism
7. The presence of which element is confirmed by the Lassaigne's test?
   a) Hydrogen b) Nitrogen c) Carbon d) Oxygen
8. What is the hybridization of carbon in ethyne (C_2H_2)?
   a) sp b) sp^2 c) sp^3 d) sp^3d
9. Which of the following is not a characteristic of homologous series?
   a) Same general formula b) Gradation in physical properties
   c) Similar chemical properties d) Different functional groups

10. What is the IUPAC name of CH_3-CH_2-OH?
   a) Methanol b) Ethanol c) Propanol d) Butanol
11. In which of the following compounds is hydrogen bonding the strongest?
   a) Methanol b) Ethanol c) Water d) Ammonia
12. Which of the following is an electrophile?
   a) OH^- b) NH_3 c) NO_2^+ d) CH_4
13. Which of the following compounds shows geometrical isomerism?
   a) Propene b) 1-Butene c) 2-Butene d) Butane
14. Which of the following is a primary alcohol?
   a) CH_3CH_2OH b) $(CH_3)_2CHOH$ c) $(CH_3)_3COH$ d) C_6H_5OH

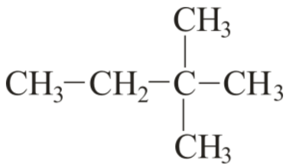
15. What is the general formula for alkanes?
   a) C_nH_{2n+2} b) C_nH_2n c) CnH_{2n-2} d) C_nH_n
16. Which of the following is a saturated hydrocarbon?
   a) Ethene b) Ethyne c) Benzene d) Propane

Answer 1B, 2C, 3C, 4B, 5D, 6B, 7B, 8A, 9D, 10B, 11C, 12C, 13C, 14A, 15A, 16D

2 Marks Questions

1. Define hybridization and give an example.
2. What are isomers? Provide an example of structural isomerism.

3. Explain the term "functional group" with an example.
4. Describe the significance of homologous series in organic chemistry.
5. How does the inductive effect influence the properties of organic compounds?
6. What is the difference between a nucleophile and an electrophile? Give an example of each.
7. What is a free radical? How is it different from an ion?
8. What are Saturated and unsaturated hydrocarbons?
9. How are alkanes prepared by Grignard's reagent?
10. Give mechanism of Wurtz reaction.
11. How will you convert acetaldehyde to ethane and acetone to propane?
12. Boiling points of isomeric alkanes goes on decreasing with increased branching. Why?
13. Alkanes with even no. of carbon atoms have high melting point as compare to alkanes with odd no. of carbon atoms why?
14. Give mechanism of sulphonation of alkanes?
15. *n*-pentane has higher boiling point than neo pentane. Explain.



16. Mention primary, secondary and tertiary carbons and hydrogens in the following compound.
17. Eclipsed conformation is less stable than staggered conformation of ethane. Explain.
18. What is geometrical isomerism and what is its cause?
19. What are the necessary conditions for the geometrical isomerisation?
20. How are alkenes prepared by Kolbe's Electrolytic process?
21. Why alkenes undergo electrophilic addition and not electrophilic substitution reaction?
22. Explain and Justify Markownikoff's rule.
23. Give mechanism of Kharash effect.
24. Give ozonolysis reaction of ethene.
25. How is structure of alkene elucidated by ozonolysis ?
26. What is lindlar's catalyst? What is its use?
27. Cis alkenes show higher boiling point as compared to trans-isomer. Why?

5 Marks Long Questions

1. Discuss the different types of structural isomerism with suitable examples.

2. Explain the concept of resonance in organic chemistry. Provide examples and discuss the conditions necessary for resonance to occur.

CHAPTER 9: HYDROCARBONS

MCQS.

1. Which represents an alkane?
   a) C_5H_8 b) C_8H_6 c) C_9H_{10} d) C_7H_{16}

2. The shape of methane molecule is?
   a) Linear b) Trigonal planar c) Square planar d) Tetrahedral

3. The alkane that yields two isomeric monobromo derivatives is?
   a) Neopentane b) Ethane c) Methane d) Propane

4. Isomers of a compound must have ?
   a) Same physical properties b) Same chemical properties
   c) Same structural properties d) Same molecular weight

5. When isomers have the same structural formula but differ in relative arrangement of atoms or groups are called?
   a) Mesomeres b) Stereo isomers c) Optical isomers d) Geometrical isomers

6. Find the correct order for relative energies of butane conformations?
   a) Staggered<skewed<eclipsed b) Skewed<staggered<eclipsed
   c) Skewed<eclipsed<staggered d) Staggered<eclipsed<skewed

7. The bond angle between H-C-C bonds in ethane is?
   a) 120 degree b) 180 degree c) 109 degree d) 109.5 degree

8. Paraffin wax is?
   a) Saturated hydrocarbon b) Unsaturated hydrocarbon
   c) Alcohol d) Ester

9. Chloroethane reacts with Na in presence of dry ether. The product is?
   a) Ethane b) Propane c) Butane d) Ethane

10. Halogenation of alkane is an example of?
    a) Electrophilic substitution b) Nucleophilic substitution
    c) Free radical substitution d) Addition reaction

11. Which represents the general formula of an alkane?
    a) C_nH_{2n+2} b) C_nH_{2n} c) C_nH_{2n-1} d) $C_{2n}H_{2n+1}$

12. Which alkane is known as marsh gas?
    a) CH_4 b) C_2H_6 c) C_3H_8 d) C_4H_{10}

13. Hydrocarbons are organic compounds with elements?
    a) Hydrogen b) Oxygen c) Carbon d) Both hydrogen and carbon

14. The step in which Cl-Cl bond hemolysis occurs is called?
    a) Initiation step b) Propagation step
    c) Intermediate step d) Termination step

15. The hybridization of carbon atoms in alkanes is?
    a) Sp b) Sp^2 c) Sp^3 d) Sp^3d

Answer

1D, 2D, 3D, 4D, 5B, 6A, 7D, 8A, 9C, 10C, 11A, 12A, 13D, 14A, 15C

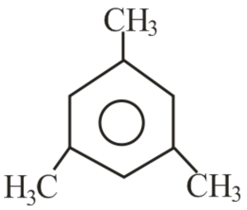
Two marks questions:

1. Write the structures of the products A and B in the following reaction:
   $HC \equiv CH \xrightarrow{Na} A \xrightarrow{CH3Br} B$
2. What effect does branching have on the boiling point of an alkane and why?
3. What is the difference between isomers and conformers?
4. Draw New man projection formula for conformations of ethane?
5. Wurtz reaction cannot be used for the preparation of unsymmetrical alkanes. Give reasons.
6. Melting point of cis-but-2-ene is lower than that of trans-but-2-ene. Give reason.
7. Draw the structures of cis and trans hex-2-ene.
8. Give a chemical test to distinguish between ethane and ethene.
9. What do you understand by peroxide effect(kharasch effect)?
10. Arrange ethane, ethene and ethyne in the order of increasing acidity.

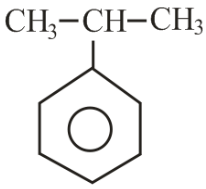
Three marks questions:

1. Define isomerism. Write all the structural isomers of hexane and arrange them in increasing order of boiling points.
2. Write short note on:
3. Wurtz reaction
4. Kolbe's electrolysis reaction
5. Ozonolysis
6. Friedel craft's acylation reaction
7. Corey house reaction
8. Alkenes show geometrical isomerism while alkanes do not. Give a suitable explanation.
9. Discuss the structure of benzene with an emphasis on resonance and orbital pictures.
10. What are aromatic hydrocarbons?
11. Give IUPAC names of the following compounds.

(i)

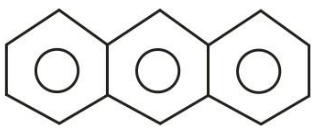


(ii)

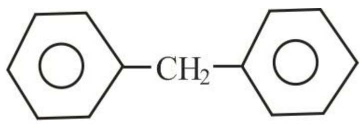


12.

(i)



(ii)



13. What is aromaticity?
14. How will you convert n-hexane to benzene?
15. How will you convert benzene to benzoic acid?
16. How will you convert benzene to benzaldehyde?
17. What are activating groups? Explain it with example.
18. Give mechanism of nitration of chlorobenzene.
19. What are electron withdrawing groups? Why are they meta-directing?
20. Give mechanism of chlorination of Nitrobenzene.
21. Give mechanism of Friedal-craft acylation reaction.
22. (i) How will you convert benzene to benzophenone?
    (i) How will you convert benzene to acetophenone?
23. Give mechanism of Sulphonation of benzene.
24. (i) Give mechanism of nitration of benzene.
    (i) How will you prepare benzene from diene's?
25. How is structure of benzene deduced? Discuss in detail.
26. Discuss evidences in favour of resonating structure of benzene.
27. Why does benzene undergo electrophilic substitutions reactions easily and nucleophilic substitutions with difficulty?
28. How would you convert following compounds into benzene?
29. Ethyne (ii) Ethene (iii) Hexane
30. Arrange benzene, n-hexane and ethyne in decreasing order of acidic behaviour. Also give reasons.
31. Out of benzene, m-dinitrobenzene and toluene which will undergo nitration most easily and why?
32. Although benzene is highly unsaturated yet it does not prefer to undergo addition reactions. Explain.
33. Why is benzene extra ordinarily stable though it contains three double bonds?
34. What are the necessary conditions for any system to be aromatic?

SAMPLE PAPER CHEMISTRY 2026

CLASS 11TH

Times: 3 Hrs **Subject- Chemistry** M.M:70

<u>Section (A) (One Mark Questions)</u>

Q1

(i) Which of the following is dependent of temperature?
a) Molarity b) Molality c) Mole fraction d) Mass Percentage

(ii) 4g of NaOH is dissolved in 100ml of solution. Molarity of the solution is:
a) 1 M b) 10 M c) 0.1 M d) 4 M

(iii) The number of moles in 32g of Methane is:
a) 3.0×10^{23} b) 12.04×10^{23} c) 6.02×10^{23} d) 2.4×10^{23}

(iv) The molecule that has a linear structure is :

a) SO_2 b) CO_2 c) NO_2 d) SiO_2

(v) Which of the following compounds have highest covalent character:

a) LiCl b) LiBr c) LiF d)LiI

(vi) In an open system for maximum work, the process must be:

a) Reversible b) Irreversible c) both d) None

(vii) For a buffer solution which of the following is true?

a) pH does not changes at all on addition of acid or base.
b) pH change is very little on addition of acid or base.
c) It is mixing of strong acid and its salt.
d) It is mixing of strong base and its salt.

(viii) $\frac{K_P}{K_C}$ for the reaction $CO_{2(g)} + \frac{1}{2}O_{2(g)} \rightleftharpoons CO_{2(g)}$ is

a) 1 b) RT c) RT/2 d) $(RT)^{-1/2}$

(ix) Oxidation involves:

a) Gain of Electron
b) Addition of Hydrogen
c) Decrease in Oxidation number
d) Loss of Electron

(x) The common name of 2-Butanone is :

a) Acetone b) Butyraldehyde
c) Acetaldehyde d) Ethyl methyl ketone

Comprehension:

The f-Block. element are called Inner transition elements because in the transition elements of d-block, the elements are filled in (n-1)d subshell while in the inner trans1tIon element of f-block, the filling of electrons takes place in (n -2)f sub shell which happens to be one inner sub shell there are two series of such elements each having fourteen elements. In the first series, the filling of electrons takes place in the 4f sub shell It is known as lanthanoid series Since it follows Lanthanum of d block and also a member of group 3 Similarly, in the second series the filling takes place In the 5f subshell It is called actinoid series as it follows actinium of d-block also belonging to group 3.

(xi) Which block elements are known as inner transition elements?
(xii) In d-block elements, the electrons are fil led in which sub shell?
(xiii) Which series is known as Lanthanoid series?
(xiv) What do you mean by transition elements?
(xv) What are actinoids?

True or False

(xvi) During isothermal expansion of an ideal gas there is no change in internal energy.
(xvii) For a spontaneous process ΔS_{system} = +ve
(xviii) Fluorine cannot have +1 oxidation state
(xix) Cis and trans isomers have different dipole moments.
(xx) Resonance brings down the stability of the molecule.

Section-B (2 Mark Question)

Q2. Calculate the percentage composition of various elements in $CaCO_3$.

OR

Calculate the mole fraction of ethanol in a sample of rectified spirit which contains 92% ethanol by mass.

Q3. State and explain Aufbau's principle?

Q4. Find the energy of photon which-

(i) Corresponds to light of frequency 3 X 10^{15} H_2
(ii) Have wavelength of 0.50 °A

Q5. Noble gasses have high ionization enthalpy in their respective period. Why?

Q6. Give differences between sigma and pi bond.

Q7. State and explain first law of thermodynamics.

Q8. Derive relationship between Cp and Cv

Q9. Write down the expression for Kp and Kc for the reaction:

$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$

Also, write relation between Kp and Kc for the given reaction.

Q10. Write formula of conjugate acids of

(i) ^{-}OH (ii) HCO_3^{-} (iii) CO_3^{-2} (iv) $HCOO^{-}$

Q11. Determine Kc for the reaction:

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

Kp= 3.4 bar^{-1} at 1000° C

OR

Calculate the pH value of 10^{-4} M HNO_3 solution.

Q12. Give difference between valency and oxidation no.

OR

Identify the oxidizing and reducing agent in the following reaction:

$2Zn_{(s)} + O_2(g) \rightarrow 2ZnO_{(s)}$

Q13. Explain tetra covalency of carbon

OR

Define geometrical isomerism with example.

Q14. Out of n-pentane and n-hexane, which has higher melting paint and why?

Q15. State and explain Markovnikov's rule with example.

Section-C (3 Marks Questions)

Q16. A compound contains carbon, hydrogen and oxygen. Gave the following analytical data.

C = 40.0% H = 6.67% O = 53.3%

Calculate the molecular formula of the compound if its molecular mass is 180.

Q17. How many photons of light with a wavelength 4000 pm are necessary to provide 1 J energy.

OR

The uncertainty in the momentum of a particle is 2.2 X 10^{-8} $kgms^{-1}$. With what accuracy can its position be determined? Planck's constant

h = 6.626 X 10^{-34} Js

Q18. Calculate the oxidation state of each sulphur atoms in the following compounds.

(i) $Na_2S_2O_3$ (ii) Na_2SO_3 (iii) Na_2SO_4

OR

What do you mean by Redox Reactions? Explain briefly.

Q19. What are carbocations? Explain the order of stability of various carbocations.

Section-D (5 Marks Questions)

Q20. A) On the basis of VSEPR theory. Discuss the shape of PCl_5 molecule.

B) Which out of NH_3 and NF_3 has higher dipole moment and why?

OR

Draw the molecular orbital diagram of F_2 molecule and write its

i) Molecular orbital electronic configuration. ii)Bond order

Q21.explain the following name reactions:

i)Friedel craft alkylation reaction ii)Ozonolysis iii)Wurtz reaction

iv)Dehydrohalogenation v)Etard reaction

OR

Write a note on sytzeff rule. Explain in detail with the help of an example.

What is peroxide effect? Also write the reaction involved.