DIRECTORATE OF EDUCATION
Govt. of NCT, Delhi

SUPPORT MATERIAL
(2025-2026)

CHEMISTRY
Class : XI

Under the Guidance of

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Secretary (Education)

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Director, Education

Dr. Rita Sharma
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पांडुरंग के. पोले, भा.प्र.से
सचिव (शिक्षा)

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SECRETARY (Education)

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MESSAGE

The Directorate of Education remains steadfast in its vision to achieve excellence in the academic domain and its commitment to develop meaningful, engaging, and child-friendly learning content.

Each year, the Directorate carefully reviews and updates the Support Material to ensure alignment with the latest CBSE guidelines and emerging academic developments.

The Support Material provides comprehensive academic support through well-structured practice questions and exercises that strengthen conceptual understanding and exam readiness and aims to nurture students' critical thinking, analytical abilities, and problem-solving skills. Through such sustained efforts, the Directorate of Education continues to guide students towards academic excellence and holistic growth.

This Support Material is intended to bridge classroom learning and examination preparation, enabling students to consolidate knowledge through systematic practice. It has been thoughtfully designed for students, with the belief that its effective use will strengthen their understanding and support them in achieving their learning goals with confidence.

I appreciates the dedication and collaborative effort of all those involved in the development of this material and extends my best wishes to all students —may this Support Material serve as an essential academic aid, enhancing students' confidence and preparedness for examinations.

Best Wishesh.

(Pandurang K. Pole)

VEDITHA REDDY, IAS
Director, Education & Sports

सत्यमेव जयते

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DE-5/228/Exam/Message/S-M/2018/402
dated - 09/05/25

MESSAGE

Education is the cornerstone of a progressive society, and providing students with the right learning resources is essential for their academic and personal growth. Keeping this in mind, the Directorate of Education, GNCT of Delhi, develops comprehensive Support Material every year for various subjects of Classes IX to XII.

The support material serves as an additional study resource to supplement textbooks by offering clear and easy-to-understand explanation of complex topics. Our dedicated team of expert faculty members has meticulously reviewed and updated this material, aligning it with the latest CBSE syllabus, question paper pattern and assessment guidelines. Our effort is to simplify difficult concepts and make them more accessible to students, helping them save time and effort with ready references for effective preparation.

As Ruskin Bond beautifully said, "Education must inspire the spirit of inquiry, Creativity and Joy" True learning goes beyond memorisation-it encourages curiosity, fosters creativity, and makes the learning process meaningful and enjoyable.

In alignment with the vision of NEP 2020, the CBSE framework now places emphases on competency-based assessments for 50% of the evaluation, highlighting the need for students to develop critical thinking and problem-solving skills. The Support Material is designed to help students analyse concepts deeply, think innovatively, and apply their knowledge affectively, ensuring they are well-prepared not only for exams but also for real-life challenges.

I appreciate the dedicated efforts of the entire team of subject experts in developing this valuable learning resource. I am confident that both teachers and students will make the best use of these material to enhance learning and academic success.

Wishing all students great success in their exam and a bright, fulfilling future ahead.

Veditha

(VEDITHA REDDY, IAS)

Dr. RITA SHARMA
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D.O. No. DE.5/228/Exam/Message/SM/2019/5to
Dated: 02/07/2025

MESSAGE

"Children are not things to be molded, but are people to be unfolded." - Jess Lair

In line with this insightful quote, the Directorate of Education, Delhi, has always made persistent efforts to nurture and unfold the inherent potential within each student. This support material is a testimony to this commitment.

The support material serves as a comprehensive tool to facilitate a deeper understanding of the curriculum. It is crafted to help students not only grasp essential concepts but also apply them effectively in their examinations. We believe that the thoughtful and intelligent utilization of these resources will significantly enhance the learning experience and academic performance of our students.

Our expert faculty members have dedicated themselves to the support material to reflect the latest CBSE guidelines and changes. This continuous effort aims to empower students with innovative approaches, fostering their problem-solving skills and critical thinking abilities.

I extend my heartfelt congratulations to the entire team for their invaluable contribution to creating a highly beneficial and practical support material. Their commitment to excellence ensures that our students are well-prepared to meet the challenges of the CBSE examinations and beyond.

Wishing you all success and fulfilment in your educational journey.

(Dr. Rita Sharma)

DIRECTORATE OF EDUCATION

Govt. of NCT, Delhi

SUPPORT MATERIAL

(2025-2026)

CHEMISTRY

Class : XI

NOT FOR SALE

PUBLISHED BY : DELHI BUREAU OF TEXTBOOKS

भारत का संविधान

उद्देशिका

हम, भारत के लोग, भारत को एक [1]**[संपूर्ण प्रभुत्व-संपन्न समाजवादी पंथनिरपेक्ष लोकतंत्रात्मक गणराज्य]** बनाने के लिए, तथा उसके समस्त नागरिकों को :

सामाजिक, आर्थिक और राजनैतिक **न्याय,**
विचार, अभिव्यक्ति, विश्वास, धर्म
और उपासना की **स्वतंत्रता,**
प्रतिष्ठा और अवसर की **समता**
प्राप्त कराने के लिए,
तथा उन **सब में**
व्यक्ति की गरिमा और [2]**[राष्ट्र की एकता**
और अखंडता] सुनिश्चित करने वाली **बंधुता**
बढ़ाने के लिए

दृढ़संकल्प होकर **अपनी इस संविधान सभा में** आज तारीख 26 नवंबर, 1949 ई. को **एतद्द्वारा इस संविधान को अंगीकृत, अधिनियमित और आत्मार्पित करते हैं।**

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1. संविधान (बयालीसवां संशोधन) अधिनियम, 1976 की धारा 2 द्वारा (3.1.1977 से) ''प्रभुत्व-संपन्न लोकतंत्रात्मक गणराज्य'' के स्थान पर प्रतिस्थापित।
2. संविधान (बयालीसवां संशोधन) अधिनियम, 1976 की धारा 2 द्वारा (3.1.1977 से) ''राष्ट्र की एकता'' के स्थान पर प्रतिस्थापित।

THE CONSTITUTION OF INDIA

PREAMBLE

WE, THE PEOPLE OF INDIA, having solemnly resolved to constitute India into a [1][**SOVEREIGN SOCIALIST SECULAR DEMOCRATIC REPUBLIC**] and to secure to all its citizens :

JUSTICE, social, economic and political;

LIBERTY of thought, expression, belief, faith and worship;

EQUALITY of status and of opportunity; and to promote among them all

FRATERNITY assuring the dignity of the individual and the [2][unity and integrity of the Nation];

IN OUR CONSTITUENT ASSEMBLY this twenty-sixth day of November, 1949 do **HEREBY ADOPT, ENACT AND GIVE TO OURSELVES THIS CONSTITUTION.**

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[1] Subs. by the Constitution (Forty-second Amendment) Act, 1976, Sec.2, for "Sovereign Democratic Republic" (w.e.f. 3.1.1977)

[2] Subs. by the Constitution (Forty-second Amendment) Act, 1976, Sec.2, for "Unity of the Nation" (w.e.f. 3.1.1977)

भारत का संविधान

भाग 4क

नागरिकों के मूल कर्तव्य

अनुच्छेद 51 क

मूल कर्तव्य – भारत के प्रत्येक नागरिक का यह कर्तव्य होगा कि वह –

(क) संविधान का पालन करे और उसके आदर्शों, संस्थाओं, राष्ट्रध्वज और राष्ट्रगान का आदर करे;

(ख) स्वतंत्रता के लिए हमारे राष्ट्रीय आंदोलन को प्रेरित करने वाले उच्च आदर्शों को हृदय में संजोए रखे और उनका पालन करे;

(ग) भारत की संप्रभुता, एकता और अखंडता की रक्षा करे और उसे अक्षुण्ण बनाए रखे;

(घ) देश की रक्षा करे और आह्वान किए जाने पर राष्ट्र की सेवा करे;

(ङ) भारत के सभी लोगों में समरसता और समान भ्रातृत्व की भावना का निर्माण करे जो धर्म, भाषा और प्रदेश या वर्ग पर आधारित सभी भेदभावों से परे हो, ऐसी प्रथाओं का त्याग करे जो महिलाओं के सम्मान के विरुद्ध हों;

(च) हमारी सामासिक संस्कृति की गौरवशाली परंपरा का महत्त्व समझे और उसका परिरक्षण करे;

(छ) प्राकृतिक पर्यावरण की, जिसके अंतर्गत वन, झील, नदी और वन्य जीव हैं, रक्षा करे और उसका संवर्धन करे तथा प्राणिमात्र के प्रति दयाभाव रखे;

(ज) वैज्ञानिक दृष्टिकोण, मानववाद और ज्ञानार्जन तथा सुधार की भावना का विकास करे;

(झ) सार्वजनिक संपत्ति को सुरक्षित रखे और हिंसा से दूर रहे;

(ञ) व्यक्तिगत और सामूहिक गतिविधियों के सभी क्षेत्रों में उत्कर्ष की ओर बढ़ने का सतत् प्रयास करे, जिससे राष्ट्र निरंतर बढ़ते हुए प्रयत्न और उपलब्धि की नई ऊँचाइयों को छू सके; और

(ट) यदि माता-पिता या संरक्षक है, छह वर्ष से चौदह वर्ष तक की आयु वाले अपने, यथास्थिति, बालक या प्रतिपाल्य को शिक्षा के अवसर प्रदान करे।

Constitution of India

Part IV A (Article 51 A)

Fundamental Duties

It shall be the duty of every citizen of India —

(a) to abide by the Constitution and respect its ideals and institutions, the National Flag and the National Anthem;

(b) to cherish and follow the noble ideals which inspired our national struggle for freedom;

(c) to uphold and protect the sovereignty, unity and integrity of India;

(d) to defend the country and render national service when called upon to do so;

(e) to promote harmony and the spirit of common brotherhood amongst all the people of India transcending religious, linguistic and regional or sectional diversities; to renounce practices derogatory to the dignity of women;

(f) to value and preserve the rich heritage of our composite culture;

(g) to protect and improve the natural environment including forests, lakes, rivers, wildlife and to have compassion for living creatures;

(h) to develop the scientific temper, humanism and the spirit of inquiry and reform;

(i) to safeguard public property and to abjure violence;

(j) to strive towards excellence in all spheres of individual and collective activity so that the nation constantly rises to higher levels of endeavour and achievement;

* (k) who is a parent or guardian, to provide opportunities for education to his child or, as the case may be, ward between the age of six and fourteen years.

Note: The Article 51A containing Fundamental Duties was inserted by the Constitution (42nd Amendment) Act, 1976 (with effect from 3 January 1977).

*(k) was inserted by the Constitution (86th Amendment) Act, 2002 (with effect from 1 April 2010).

Support Material Preparation Team

Class XI – Chemistry

S.No.	Subject Expert	Designation	School
1.	Ashok Kumar	Vice Principal	SBS AFPS SOSE Jharoda Kalan
2.	Sandeep	PGT (Chem.)	GBSSS Alipur
3.	Amit Kumar	PGT (Chem.)	S.V. Sec. 8, Rohini
4.	Manjeet Singh Hooda	PGT (Chem.)	SBV Keshav Puram
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CHEMISTRY (CODE NO. 043)

Rationale

Higher Secondary is the most crucial stage of school education because at this juncture specialized discipline based, content -oriented courses are introduced. Students reach this stage after 10 years of general education and opt for Chemistry with a purpose of pursuing their career in basic sciences or professional courses like medicine, engineering, technology and study courses in applied areas of science and technology at tertiary level. Therefore, there is a need to provide learners with sufficient conceptual background of Chemistry, which will make them competent to meet the challenges of academic and professional courses after the senior secondary stage.

The new and updated curriculum is based on disciplinary approach with rigour and depth taking care that the syllabus is not heavy and at the same time it is comparable to the international level. The knowledge related to the subject of Chemistry has undergone tremendous changes during the past one decade. Many new areas like synthetic materials, bio -molecules, natural resources, industrial chemistry are coming in a big way and deserve to be an integral part of chemistry syllabus at senior secondary stage. At international level, new formulations and nomenclature of elements and compounds, symbols and units of physical quantities floated by scientific bodies like IUPAC and CGPM are of immense importance and need to be incorporated in the updated syllabus. The revised syllabus takes care of all these aspects. Greater emphasis has been laid on use of new nomenclature, symbols and formulations, teaching of fundamental concepts, application of concepts in chemistry to industry/ technology, logical sequencing of units, removal of obsolete content and repetition, etc.

Objectives

The curriculum of Chemistry at Senior Secondary Stage aims to:

- promote understanding of basic facts and concepts in chemistry while retaining the excitement of chemistry.
- make students capable of studying chemistry in academic and professional courses (such as medicine, engineering, technology) at tertiary level.
- expose the students to various emerging new areas of chemistry and apprise them with their relevance in future studies and their application in various spheres of chemical sciences and technology.
- equip students to face various challenges related to health, nutrition, environment, population, weather, industries and agriculture.
- develop problem solving skills in students.

- expose the students to different processes used in industries and their technological applications.
- apprise students with interface of chemistry with other disciplines of science such as physics, biology, geology, engineering etc.
- acquaint students with different aspects of chemistry used in daily life.
- develop an interest in students to study chemistry as a discipline.
- integrate life skills and values in the context of chemistry.

Course Structure

Class : XI (Theory) (2023-24)

Chemistry

Total period (Theory 160 + Practical 60)

Time : 3 Hours — Total Marks : 70

Unit I	Some Basic Concepts of Chemistry	12	07
Unit II	Structure of Atom	14	09
Unit III	Classification of Elements and Periodicity in Properties	08	06
Unit IV	Chemical Bonding and Molecular Structure	14	07
Unit V	Chemical Thermodynamics	16	09
Unit VI	Equilibrium	14	07
Unit VII	Redox Reactions	06	04
Unit VIII	Organic Chemistry: Some Basic Principles and Techniques	14	11
Unit IX	Hydrocarbons	12	10
	Total	**110**	**70**

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

General Introduction: Importance and scope of chemistry.

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

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

Unit II : Structure of Atom 14 Periods

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

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

Significance of classification, Brief History of the Development of periodic table, Modern periodic law and the present form of periodic table, periodic trends in properties of elements -atomic radii, ionic radii, inert gas radii, Ionization enthalpy, electron gain enthalpy, electronegativity, valency. Nomenclature of elements with atomic number greater than 100.

Unit IV : Chemical Bonding and Molecular Structure 14 Periods

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

Unit VI : Chemical Thermodynamics 16 Periods

Concepts of System and types of systems, surroundings, work, heat, energy, extensive and intensive properties, state functions. First law of thermodynamics -internal energy and enthalpy, heat capacity and specific heat, measurement of ΔU and ΔH, Hess's law of constant heat summation, enthalpy of bond dissociation, combustion, formation, atomization, sublimation, phase transition, ionization, solution and dilution. Second law of Thermodynamics (brief introduction). Introduction of entropy as a state function, Gibb's energy change for spontaneous and non- spontaneous processes, criteria for equilibrium. Third law of thermodynamics (brief introduction).

Unit VII : Equilibrium 14 Periods

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

Unit VIII: Redox Reactions 06 Periods

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

Unit XII : Organic Chemistry -Some Basic Principles and Technique 14 Periods

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

Unit XIII: Hydrocarbons 12 Periods

Classification of Hydrocarbons Aliphatic Hydrocarbons:

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

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

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

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

PRACTICALS

Evaluation Scheme for Examination	**Marks**
Volumetric Analysis	08
Salt Analysis	08
Content Based Experiment	06
Project Work	04
Class record and viva	04
Total	**30**

PRACTICAL SYLLABUS **Total Periods 60**

Micro-chemical methods are available for several of the practical experiments. Wherever possible such techniques should be used :

A. Basic Laboratory Techniques

1. Cutting glass tube and glass rod
2. Bending a glass tube
3. Drawing out a glass jet
4. Boring a cork

B. Characterization and Purification of Chemical Substances

1. Determination of melting point of an organic compound.
2. Determination of boiling point of an organic compound
3. Crystallization of impure sample of any one of the following: Alum, Copper Sulphate, Benzoic Acid.

C. Experiments based on pH

(a) Any one of the following experiments :

- Determination of pH of some solutions obtained from fruit juices, solution of known and varied concentrations of acids, bases and salts using pH paper or universal indicator.
- Comparing the pH of solutions of strong and weak acids of same concentration.
- Study the pH change in the titration of a strong base using universal indicator.

(b) Study the pH change by common-ion in case of weak acid and weak bases.

D. Chemical Equilibrium.

One of the following experiments:

(a) Study the shift in equilibrium between ferric ions and thiocyanate ions by increasing/decreasing the concentration of either of the ions.

(b) Study the shift in equilibrium between $[Co(H_2O)_6]^{2+}$ and chloride ions by changing the concentration of either of the ions.

E. Quantitative Estimation

(i) Using a mechanical balance/electronic balance

(ii) Preparation of standard solution of Oxalic acid.

(iii) Determination of strength of a given solution of Sodium hydroxide by titrating it against standard solution of Oxalic acid.

(iv) Preparation of standard solution of Sodium carbonate.

(v) Determination of strength of a given solution of hydrochloric acid by titrating it against standard Sodium Carbonate solution.

F. Qualitative Analysis

(a) Determination of one anion and one cation in a given salt

Cations- Pb^{2+}, Cu^{2+}, As^{3+}, Al^{3+}, Fe^{3+}, Mn^{2+}, Ni^{2+}, Zn^{2+}, Co^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, Mg^{2+}, NH^{4+}

Anions – CO_3^{2-} , S^{2-}, NO^{2-} , SO_3^{2-}, SO_4^{2-}, NO_3^- , Cl^- , Br^-, I^-, PO_4^{3-}, CH_3COO^-

(**Note:** Insoluble salts excluded)

(b) Detection of -Nitrogen, Sulphur, Chlorine in organic compounds.

(c) Projects

Scientific investigations involving laboratory testing and collecting information from other sources.

A few suggested Projects

- Checking the bacterial contamination in drinking water by testing sulphide ion
- Study of the methods of purification of water
- Testing the hardness, presence of Iron, Fluoride, Chloride, etc., depending upon the regional variation in drinking water and study of causes of presence of these ions above permissible limit (if any).

- Investigation of the foaming capacity of different washing soaps and the effect of addition of Sodium carbonate on it
- Study the acidity of different samples of tea leaves.
- Determination of the rate of evaporation of different liquids Study the effect of acids and bases on the tensile strength of fibers.
- Study of acidity of fruit and vegetable juices.

Note: Any other investigatory project, which involves about 10 periods of work, can be chosen with the approval of the teacher.

Practical Examination for Visually Impaired Students
Class XI

Note: Same Evaluation scheme and general guidelines for visually impaired students as given for Class XII may be followed.

A. List of apparatus for identification for assessment in practicals (All experiments)

Beaker, tripod stand, wire gauze, glass rod, funnel, filter paper, Bunsen burner, test tube, test tube stand, dropper, test tube holder, ignition tube, china dish, tongs, standard flask, pipette, burette, conical flask, clamp stand, dropper, wash bottle

- Odour detection in qualitative analysis
- Procedure/Setup of the apparatus

B. List of Experiments

(a) Characterization and Purification of Chemical Substances

1. Crystallization of an impure sample of any one of the following: copper sulphate, benzoic acid

(b) Experiments based on pH

1. Determination of pH of some solutions obtained from fruit juices, solutions of known and varied concentrations of acids, bases and salts using pH paper
2. Comparing the pH of solutions of strong and weak acids of same concentration.

C. Chemical Equilibrium

1. Study the shift in equilibrium between ferric ions and thiocyanate ions by increasing/decreasing the concentration of either ion.
2. Study the shift in equilibrium between $[Co(H_2O)_6]^{2+}$ and chloride ions by changing the concentration of either of the ions.

D. Quantitative estimation

1. Preparation of standard solution of oxalic acid.
2. Determination of molarity of a given solution of sodium hydroxide by titrating it against standard solution of oxalic acid.

E. Qualitative Analysis

1. Determination of one anion and one cation in a given salt
2. Cations – NH^+_4

   Anions – $(CO_3)^{2-}$, S^{2-}, $(SO_3)^{2-}$, Cl^-, CH_3COO^-

   (**Note:** insoluble salts excluded)
3. Detection of Nitrogen in the given organic compound.
4. Detection of Halogen in the given organic compound.

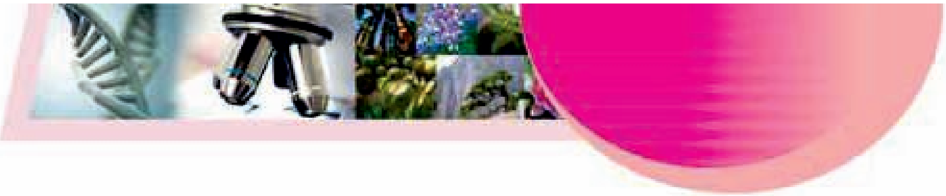
   Note: The above practicals may be carried out in an experiential manner rather than recording observations.

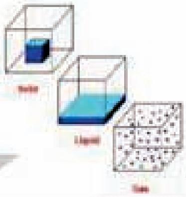
Prescribed Books:

1. Chemistry Part – I, Class-XI, Published by NCERT.
2. Chemistry Part – II, Class-XI, Published by NCERT,

Chemistry - XI

Index





Chapter - 1

Some Basic Concepts of Chemistry

FAST TRACK : QUICK REVISION

- **Matter :** Anything that has mass and occupies space.
- **Precision :** It refers to the closeness of various measurements for the same quantity.
- **Accuracy :** It refers to the agreement of a particular value to the true value of the result.
- **Mass and weight :** Mass of a substance is the amount of matter present in body, while weight is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.
- **Volume :** $1\ L = 1\ dm^3 = 10^3 cm^3 = 10^{-3}\ m^3$
- **Temperature :** $K = °C + 273.15$; $\frac{°F - 32}{9} = \frac{°C}{5}$
- **Standard Temperature Pressure (STP) :** 0°C (273.15 K) temperature and 1 atm pressure.
- **Normal Temperature Pressure (NTP) :** 20°C (293.15 K) temperature and 1 atm pressure.
- **Standard Ambient Temperature Pressure (SATP) :** 25°C (298.15 K) temperature and 1 atm pressure
- **Scientific Notation :** Expressing a number in the form $N \times 10^n$, and N can vary between 1 to 9.99.
- **Significant figures :** These are meaningful digits which are known with certainty.
- **Laws of Chemical Combination :**
  - **Law of Conservation of Mass (Antonie Lavoisier) :** Mass can neither be created nor be destroyed.
  - **Law of Definite Proportions (Joseph Proust) :** A given compound always contains the same elements in the same proportion by mass.

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

  e.g., $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

  2 Vol 1 Vol 2 Vol

  (at same T, P)

- **Atomic Mass :** It is defined as the average relative mass of an atom of an element as compared to the mass of an atom of carbon – 12 taken as 12.

  Atomic mass is represented by '*u*' (unified mass).

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

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

  For example : Molecular mass of $CH_4 = (1 \times 12) + (4 \times 1) = 16$ u

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

  Avogadro number $(N_A) = 6.022 \times 10^{23}$ mol^{-1}

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

  For example : Molar mass of $CH_4 = (1 \times 12) + (4 \times 1) = 16$g mol^{-1}

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

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

  e.g., 1 mole of CH_4 = 16g of CH_4 = 6.022×10^{23} molecules of CH_4 = 22.4L at STP

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

- **Molar Volume (V_m) :** It is volume occupied by one mole of gas at STP.

  Molar volume of a gas = 22.4L at STP (273 K, 1atm) or 22.7L at STP (273 K, 1 bar)

Calculating Molar Volume: $PV = nRT$

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

Or

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

- Combined gas equation $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- **Percentage Composition :** Mass % of the element

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

- **Empirical Formula :** It represents the simplest whole number ratio of various atoms present in a compound. For *e.g.*, CH is the empirical formula of benzene.
- **Molecular Formula** : It shows the exact number of different of atoms present in a molecule of a compound. For *e.g.*, C_6H_6 is the molecular formula of benzene.
- **Relationship between empirical and molecular formulae :** Molecular formula = n × Empirical formula

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

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

g)

(i)	1 molecule of N_2	+	3 molecules of H_2	→	2 molecules of NH_3
(ii)	1 mole of N_2	+	3 mole of H_2	→	2 mole of NH_3
(iii)	1 × 28g of N_2	+	3 × 2 g of H_2	→	2 × 17 g of NH_3
(iv)	1 × 22.4L of N_2 at STP	+	3 × 22.4L of H_2 at STP	→	2 × 22.4L of NH_3 at STP

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

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

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

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

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

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

Molarity equation : $M_1V_1 = M_2V_2$

(Before dilution) (After Dilution)

Molarity of a solution decreases on increasing temperature.

Molarity of pure water is 55.56 mol L^{-1}

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

$$\text{Molality} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

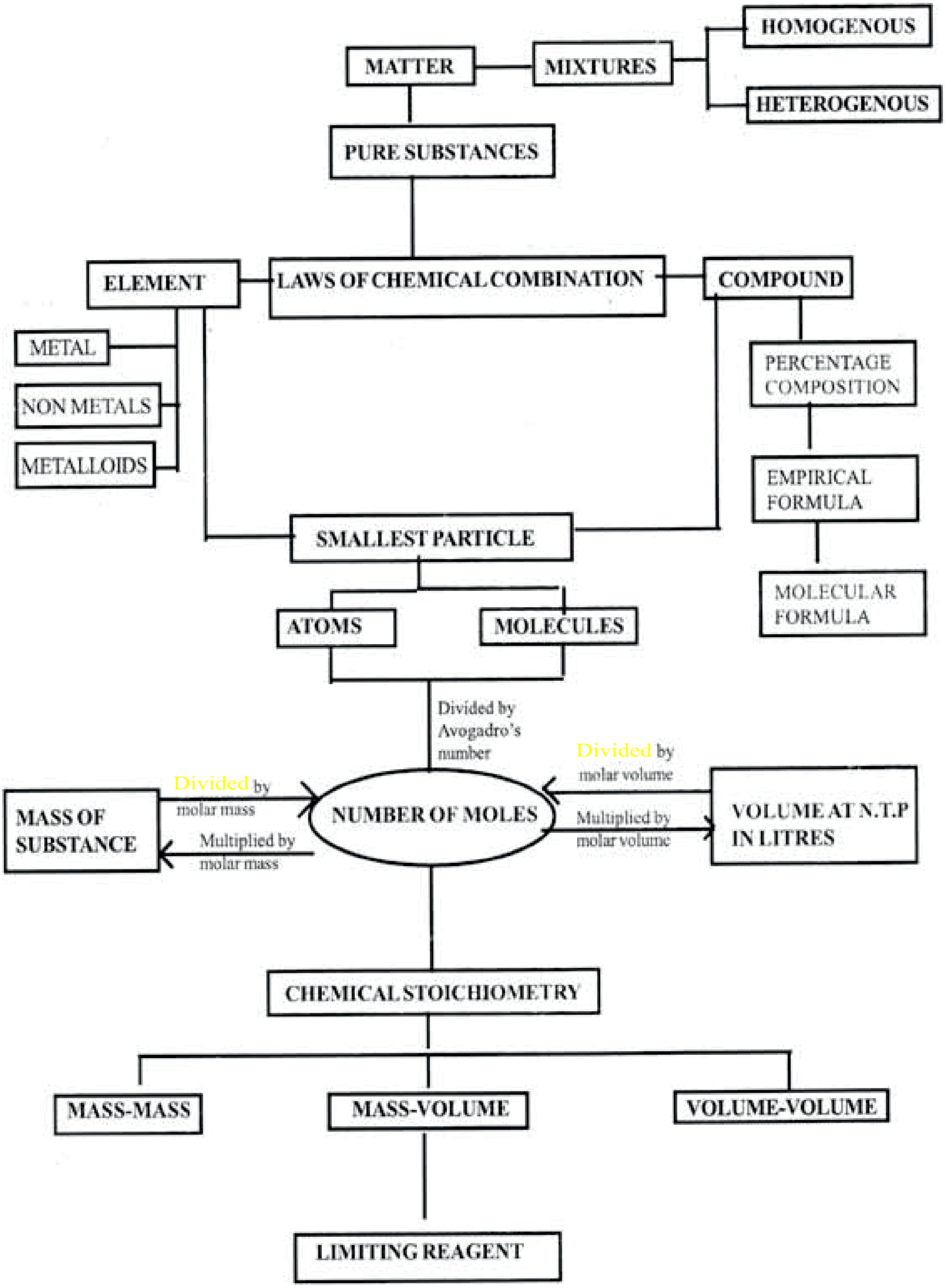
Molality is independent of temperature.

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

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

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

MIND MAP
SOME BASIC OF CONCEPTS OF CHEMISTRY



CASE BASE : QUESTIONS

1. Read the passage given below and answer the following questions:

The ideas underlying our modern understanding of thermodynamics and kinetic theory were developed during the nineteenth century. Central to these developments was the discovery that matter reacting chemically does not do so simply between equal masses of the samples involved. We now call the study of this phenomenon 'stoichiometry', defined as: 'the relationship between the amounts of substance that react together, and the products that are formed'.

Another development during the nineteenth century that was central to our modern understanding of the chemical nature of matter was the observation by Avogadro that 'equal volumes of ideal or perfect gases, at the same temperature and pressure, contain the same number of particles, or molecules'. This is now known as Avogadro's law. It provides the motivation to formulate expressions for the quantity of a sample that reacts with another sample. The most notable example of such a formulation is the gram-molecule, which has been used to refer to both a unit and a quantity.

(Reference: Milton Martin J. T. 2011A new definition for the mole based on the Avogadro constant: a journey from physics to chemistry *Phil. Trans. R. Soc.* A.3693993–4003)

The following questions are multiple choice questions. Choose the most appropriate answer:

I. The concept of stoichiometry mentioned in the study is based on the

a. formation of chemical bonds.

b. amount of reactant and product involved in a chemical reaction.

c. idea of temperature and pressure required for the reaction to occur.

d. oxidation states of reactant and product involved.

II. How much gram-molecules of H_2O are produced on combustion of 32 g of methane in excess oxygen?

a. 72 b. 4

c. 2 d. 36

III. When an antacid tablet is used, $Ca(OH)_2$ reacts with HCl in the stomach to form inert $CaCl_2$ and H_2O. If the molar mass of $Ca(OH)_2$ is 75 g/mol, how many moles of HCl are required to fully react with 150 g of $Ca(OH)_2$?

a. 4
b. 1
c. 8
d. 2

IV. What must be held constant when applying Avogadro's law?

a. pressure and temperature
b. volume and temperature
c. moles and temperature
d. pressure and volume

ANS:- I-B, II-B, III-A, IV-A

2. **Read the passage given below and answer the following questions:**

The goal of this study was to examine the means used by textbook authors to introduce, define, and explain the mole concept in high school and introductory college chemistry textbooks. The analysis was framed by four questions:

1. How is the mole defined?
2. What concepts about the atom are introduced prior to the mole?
3. Is Avogadro's constant presented as an experimentally determined value?
4. What is the context for introducing the mole?

Twenty nine high school and introductory college level chemistry texts were examined. After independent reading of appropriate sections of each text, discussion of differences, second or third readings of texts, and subsequent discussions, both authors reach 100% agreement concerning the results. Major conclusions were

1. Two ways of defining the mole dominate the texts. One way defines the mole as Avogadro's number (6.02×10^{23}) particles; the other method defines the mole in terms of carbon 12.
2. All texts that present a definition in terms of C-12 introduce and define concepts about the atom prior to introducing the mole.
3. Most texts at all levels point out that the value 6.02×10^{23} is an experimentally determined quantity.

4. Nearly all texts discuss the mole in relation to die problem of finding a way to count particles that are too small to be directly weighed. Most texts also use a familiar counting unit, such as the dozen, to introduce the mole by analogy.

(Reference: John R. Staver, Andrew T. Lumpe, A content analysis of the presentation of the mole concept in chemistry textbooks, Journal of Research in Science Teaching).

In these questions (Q. No. (i) to (iv) , a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

a) Assertion and reason both are correct statements and reason is correct explanation for assertion.

b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

c) Assertion is correct statement but reason is wrong statement.

d) Assertion is wrong statement but reason is correct statement

i. ASSERTION: Equal moles of different substances contain same number of constituent particles.

REASON: Equal weights of different substances contain the same number of constituent particles.

ii. ASSERTION: Both 106g of sodium carbonate and 12g of carbon have same number of carbon atoms.

REASON: Both contain 1 g-atom of carbon which contains 6.02×10^{23} carbon atoms

iii. ASSERTION: Both 32g SO_2 and 8g CH_4 have same number of molecules.

REASON: Equal moles of substances have equal number of molecules.

iv. ASSERTION: Average atomic mass of an element depends mainly on the heavier isotope.

REASON: The average atomic mass of an element is the sum of the masses of its isotopes, each multiplied by its natural abundance

ANS:- I-C, II-A, III-A, IV-D

MULTIPLE CHOICE QUESTIONS (MCQ)

1. Which of the following is dependent of temperature ?

(a) Molarity
(b) Molality
(c) Mole fration
(d) Mass percentage

2. 4 g of NaOH dissolved in 100 ml solution. Molarity of the solution is

(a) 1 M
(b) 10 M
(c) 0.1 M
(d) 4 M

3. Which has the maximum number of molecules among the following ?

(a) 44g of CO_2
(b) 44g of O_2
(c) 8g of H_2
(d) 64g of SO_2

4. 10 mol of Zn react with 10 mol of HCl. Calculate the number of moles of H_2 produced.

(a) 5 mol
(b) 10 mol
(c) 20 mol
(d) 2.5 mol

5. The number of oxygen atoms in 4.4g of CO_2 is approximately

(a) 1.2×10^{23}
(b) 6×10^{22}
(c) 6×10^{23}
(d) 12×10^{23}

6. The molarity of a solution obtained by mixing 750 mL of 0.5 M HCl with 250 ml of 2 M HCl will be

(a) 0.975 M
(b) 0.875 M
(c) 1.00 M
(d) 1.175 M

7. Number of atoms of He in 100 u of He (Atomic mass of He is 4 u)

(a) 25
(b) 50
(c) 100
(d) 400

8. 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of the solution is

(a) 0.02 M
(b) 0.01 M
(c) 0.001 M
(d) 0.1 M

9. A gaseous hydrocarbons gives upon combustion, 0.72 g of water and 3.08 g of CO_2. The empirical formula of the hydrocarbon is :

(a) C_6H_5 (b) C_7H_8

(c) C_2H_4 (d) C_3H_4

10. The density of solution prepared by dissolving 120 g of urea (Mol. mass = 60 u) in 1000 g of water is 1.15 g/mL. The molarity of the solution is

(a) 0.50 M (b) 1.78 M

(c) 1.02 M (d) 2.05 M

Ans: 1. (a), 2. (a), 3. (c), 4. (a), 5. (a), 6. (b), 7. (a), 8. (b),

9. (b), 10. (d)

FILL IN THE BLANKS

1. 17 g of NH_3 gas will occupy a volume of ______________ cm^3 at STP.
2. The number of Li atoms in ______________ g. is 6.022×10^{24} atoms.
3. (1/12)th of the mass of carbon atom is ______________
4. Number of atoms of oxygen in 24 g of O_3 is ______________
5. The number of moles of barium carbonate which contains 1.5 moles of oxygen atoms is ______________
6. A mixture having 2 g of H_2 and 32 g of oxygen occupies a volume of ______________ at STP.
7. If the phosphate of a metal has the formula MPO_4 the formula of the metallic sulphate is ______________
8. At NTP, the mass of 1 litre of gas is 3 g. Molecular mass of the gas is ______________
9. The percentage mass of magnesium in chlorophyll is 2.68% The number of magnesium atoms in 2 g of chlorophyll is ______________
10. The mass of one molecule of carbon dioxide is ______________
11 Percentage of nitrogen in urea is ____________
12. Number of carbon atoms present in 18 g of glucose ($C_6H_{12}O_6$)

13. 0.5 mole of triatomic gas contains ____________________atoms.

14. A binary compound contains 50% A (at. mass = 16) and 50% B (at. mass 32). The empirical formula of the compound is _____________.

15. The number of hydrogen atoms in 60 u of ethane is____________

Ans: 1. 22400 2. 70 g 3. 1 u
4. 9.033×10^{23} 5. 0.5 6. 44.8 litre
7. $M_2(SO_4)_3$ 8 67.2 9. 1.34×10^{21}
10. 7.3×10^{-23} 11. 46.67 12. 3.61×10^{23}
13. 9.033×10^{23} 14. A_2B 15. 7.226×10^{24}

TRUE AND FALSE TYPE QUESTIONS

Write true or false for the following statements

1. Equal volumes of different gases under similar conditions of temperature and pressure contain equal number of molecules.
2. 1 mole of $C_{12}H_{22}O_{11}$ contain 22 hydrogen atoms.
3. Nitrogen forms five oxides. It proves the law of multiple proportions.
4. The atomicity of phosphorus is four.
5 Molarity change with change in temp.
6. Empirical formula = $(\text{Molecular formula})_n$.
7. Gram-atomic mass of an element may be defnined as the mass of Avogadro's number of atoms.
8. Gay-Lussac's law of chemical combination is valid for all substances.
9. Avogadro's number varies with temperature and pressure.
10. 18 g of water vapour and 18 g of ice will contain the same number of molecules.

Ans: 1. (T) 2. (F) 3. (T) 4. (T) 5. (T)
6. (F) 7. (T) 8. (F) 9. (F) 10. (T)

MATCH THE COLUMNS

1.

Column X	Column Y	Column Z
a. 8 g CH_4	i. 0.1 mol	p. Emp. formulation → CH_2O
b. 1.7 g NH_3	ii. 0.5 mol	q. 50% oxygen
c. CH_3OH	iii. 40% carbon	r. 1.806×10^{23} atoms of hydrogen
d. $C_6H_{12}O_6$	iv. Vapour density = 16	s. 25% hydrogen

2.

Column X	Column Y
a. Molarity	i. For very dilute solution
b. Molality	ii. No units
c. mole fraction	iii. Mol L^{-1}
d. ppm	iv. independent of temperature

3.

Column X	Column Y	Column Z
a. 40 g of He	i. 3.011×10^{23} atoms	p. 0.5 moles
b. 35 g of Li	ii. 10 atoms	q. 1.67×10^{-23}
c. 40 u of He	iii. 6.022×10^{24} atoms	r. 10 moles
d. 16 g of O_2	iv. 3.011×10^{24} atoms	s. 5 moles

4.

Column X	Column Y
a. Petrol	i. Heterogenous mixture
b. Graphite	ii. Compound
c. Sucrose	iii. Element
d. Milk	iv. Homogeneous mixture

Ans: 1. a.(ii). (s), b.(i). (r), c.(iv). (q), d.(iii). (p)

2. a.(iii), b.(iv), c.(ii), d.(i)

3. a.(iii). (r), b.(iv). (s), c.(ii). (q), d.(i). (p)

4. a.(iv), b.(iii), c.(ii), d.(i)

ASSERTION AND REASON TYPE QUESTIONS

Directions for Q. No.1-10

A Both Assertion & Reason are true and the reason is the correct explanation of the assertion.

B Both Assertion & Reason are true but the reason is not the correct explanation of the assertion.

C Assertion is true statement but Reason is false.

D Assertion is false but Reason is true.

1. Assertion : A solution of table salt in a glass of water is homogeneous
Reason : A solution having same composition throughout is heterogeneous

2. Assertion : The molecular weight of oxygen is 32 amu.
Reason : The atomic weight of oxygen is 16 amu

3. Assertion : No of moles of H_2 in 0.224 L of hydrogen is 0.01 mole.
Reason : 22.4 L of H_2 at STP contain 6.023×10^{23} moles.

4. Assertion : Atomic mass of Na is 23.
Reason : An atom of sodium is 23 times heavier than 1/12th mass of C-12 isotope.

5. Assertion : Number of atoms of He in 60 u of He is 15.
Reason : Atomic weight of He is 4 u.

6. Assertion : In a gaseous reaction, the ratio of volumes of reactants and products is in agreement with their molar ratio.
Reason : Volume of gas is inversely proportional to its moles at particular temperature and pressure.

7. Assertion : The Empirical mass of ethane is half of its molecular mass.
Reason : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

8. Assertion : Significant figures for 0.200 is 3 whereas for 200 is 1
Reason : Zero at the end or right of a number are significant provided they are not on the right side of the decimal point.

9. Assertion : One molar aqueous solution has always higher concentration than One molal aqueous solution
Reason : The molality of solution depends upon the density of solution whereas molarity does not.

10. Assertion : In a combustion reaction in the air, oxygen is the limiting reagent

Reason : Limiting reagent is the reactant in a chemical reaction that limits the amount of product that can be formed.

Ans: 1.C 2.A 3.C 4.A 5.A 6. C 7.A 8.C 9.B 10.D

ONE WORD ANSWER TYPE QUESTIONS

1. What is the SI unit of density?
2. What is the SI unit of molarity?
3. Calculate the number of atoms in 32 u of He. [**Ans.** : 8]
4. What is the volume of 17 g of NH_3 gas at STP? [**Ans.** : 22.4 L]
5. How many molecules of SO_2 are present in 11.2 L at STP? [**Ans.** : 3.011×10^{23}]
6. Which has more number of atoms ? 1.0 g Na or 1.0 g Mg [**Ans.** : 1.0 g Na]
7. How many oxygen atoms are present in 16 g of ozone (O_3)? [**Ans.** : 2.007×10^{23}]
8. Calculate the number of molecules present in 22.0 g of CO_2. [**Ans.** : 3.011×10^{23}]
9. A substance has molecular formula $C_6H_{12}O_6$. What is its empirical formula.
10. Empirical formula of a compound X (Molar mass = 78 mol^{-1}) is CH. Write its molecular formula.

1-MARK QUESTIONS

1. Classify following as pure substances and mixtures : air, glucose, gold, sodium and milk.
2. Which measurement is more precise 4.0g or 4.00g ? [**Ans.** 4.00 g]

3. How many significant figures are there in (i) 3.070 and (ii) 0.0025 ?

[**Ans.** (i) 4 (ii) 2]

4. Express the following in the scientific notation : (i) 0.0048 (ii) 234,000

5. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how much volume of water vapour would be produced ?

[**Ans.** 10 volumes]

6. At STP, what will be the volume of 6.022×10^{23} molecules of H_2 ?

[**Ans.** 22.4L]

7. 1L of a gas at STP weighs 1.97g. What is molecular mass ?

[**Ans.** 44.128 g mol^{-1}]

8. Write the relationship between empirical formula and molecular formula.

9. Which is more informative ? Empirical formula or Molecular formula.

10. How are 0.5 mol Na_2CO_3 and 0.5 M Na_2CO_3 different from each other ?

11. Why molality is preferred over molarity of a solution ?

12. Define molarity of a solution.

13. What is the effect of temperature on molarity of solution ?

14. What is limiting reactant in a reaction ?

2-MARKS QUESTIONS

1. Classify following substances as element, compounds and mixtures : water, tea, silver, steel, carbon dioxide and platinum.

2. The body temperature of a normal healthy person is 37°C. Calculate its value in°F.

3. Convert 5L into m^3.

4. What does the following prefixes stand for :

(a) pico (b) nano (c) micro (d) deci

5. How many significant figures are present in the following :

(i) 4.00005

(ii) 0.004

6. Convert '450 pm' into SI unit and write the answer in scientific notation upto 2 significant figures.

[**Ans**. 4.5×10^{-10} m]

7. Hydrogen peroxide and water contain 5.93% and 11.2 % of hydrogen respectively. Show that the data illustrate law of multiple proportions.

8. The density (in g mL^{-1}) of a 3.60 M sulphuric acid solution that is 29% H_2SO_4 (Molar mass = 98 g mol^{-1}) by mass will be ....

[**Ans.** 1.21 g/mL]

9 The cost of table salt (NaCl) is Rs. 10 per Kg. calculate its cost per mole. (Molar mass of NaCl is 58.5 $gmol^{-1}$) [**Ans.** 0.58 Rs]

10 Calculate the mole fraction of the solute in a 1.00 molal aqueous solution. [**Ans.** 0.0177]

11 Dissolving 120 g of urea (Molar mass of urea = 60 g mol^{-1}) in 1000 of water gave a solution of density 1.15 g/mL. Calculate the molarity of the solution. [**Ans.** 2.05 M]

12 Calculate the percentage of N in urea. (Molar mass of urea = 60 g mol^{-1}) [**Ans.** 46.66]

13 25 ml of 3.0 M HCl are mixed with 75 mL 0f 4.0 M HCl. If the volumes are additive, the molarity of the final mixture will be. [**Ans.** 3.75 M]

14 How many atoms and molecules are present in 124 gm of phosphorus (P_4) [**Ans.** Atoms = 4 N_A & Molecules = N_A]

15 45.4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed.

The reaction is given below : $2N_2 (g) + O_2 (g) \longrightarrow 2N_2O (g)$

Which law is being obeyed in this experiment? Write the statement of the law.

16 Give one example each of a molecule in which empirical formula and molecular formula is

(i) Same (ii) Different.

17 Calculate the number of moles in the following masses :

(i) 7.85g of Fe;
(ii) 7.9mg of Ca

18 Calculate the percent of carbon, hydrogen and oxygen in ethanol (C_2H_5OH) [**Ans.** 52.14%, 13.13%, 34.73%]

19 How much copper can be obtained from 100 g of $CuSO_4$? [**Ans.** 39.8g]

20 Calculate the amount of water (g) produced by the combustion of 16 g of methane. [**Ans.** 36g]

21 How many moles of methane are required to produce 22 g CO_2 (g) after combustion? [**Ans.** 0.5 mol]

22 A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute. [Ans. 10%]

23 Calculate molarity of water if its density is 1.00 g mL^{-1}.

[**Ans.** 55.56 M]

24 Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution. [**Ans.** 0.4 M]

25 The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate molality of the solution. [**Ans.** 2.8m]

26 NH_3 gas can be prepared by Haber's process as, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. At a particular moment concentration of all the species is 2 moles; calculate the concentration of N_2 and H_2 taken initially.

[**Ans.** 3 mole, 5 moles]

3-MARKS QUESTIONS

1. Calculate the average atomic mass of Mg using the following data:

	% Natural Abudance	Molar mass
^{24}Mg	80	24
^{25}Mg	10	25
^{26}Mg	10	26

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

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

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

3. Calculate :

(i) Mass in gram of 5.8 mol N_2O

(ii) Number of moles in 8.0 g of O_2

(iii) Molar mass if 11.2 L at STP weigh 8.5 g.

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

4. In three moles of ethane (C_2H_6), calculate the following :

(i) Number of moles of carbon atom,

(ii) Number of moles of hydrogen atoms,

(iii) Number of molecules of ethane.

[**Ans.** (i) 6 moles, (ii) 18 moles, (iii) 1.81×10^{24}]

5. 16 g of an ideal gas SO_x occupies 5.6 L at STP. What is its molecular mass ? What is the value of X ? [**Ans.** 64u, x = 2]

6. Calculate the number of moles :

(i) 5.0 L of 0.75 M Na_2CO_3

(ii) 7.85 g of Fe

(iii) 34.2 g of sucrose ($C_{12}H_{22}O_{11}$)

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

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

(i) 52 moles of Ar. (ii) 52u of He (iii) 52g of He.

[**Ans.** (i) 3.13×10^{25} (ii) 13 (iii) 7.83×10^{24}]

8. Vitamin C is essential for the prevention of scurvy. Combustion of 0.2000g of vitamin C gives 0.2998g of CO_2 and 0.0819g of H_2O. What is the empirical formula of vitamin C ? [**Ans.** $C_3H_4O_3$]

9. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas? [**Ans.** CH_2C1, $C_2H_4Cl_2$]

10. A compound made up of two elements A and B has A = 70%, B = 30%. Their relative number of moles in the compound is 1.25 and 1.88, calculate :

(i) Atomic masses of the elements A and B

(ii) Molecular formula of the compound , if its molecular mass is found to be 160. [**Ans.** (i) 56 and 16, (ii) A_2B_3]

11. The reaction $2C + O_2 \longrightarrow 2CO$ is carried out by taking 24.0 g of carbon and 96.0 g of O_2. Find out.

(i) Which reactant is left in excess ?

(ii) How much of it is left ?

(iii) How many grams of the other reactant should be taken so that nothing is left at the end of the reaction ? [**Ans.** (i) O_2, (ii) 64 g, (iii) 72]

12. A 10 g sample of a mixture of calcium chloride and sodium chloride is treated with Na_2CO_3 to precipitate calcium as calcium carbonate. This $CaCO_3$ is heated to convert all the calcium to CaO and the final mass of CaO is 1.62 g. Calculate % by mass of NaCl in original solution.

[**Ans.** 67.9%]

13. 3.0 g of H_2 react with 29.0 g of O_2 yield H_2O.

(i) Which is the limiting reagent.

(ii) Calculate the maximum amount of H_2O that can be formed

(iii) Calculate the amount of reactant left unreacted

[**Ans.** H_2 , 26.8g H_2O & 5.2 g O_2]

14 Zinc and hydrochloric acid react according to the reaction:

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

If 0.30 mol Zn are added to hydrochloric acid containing 0.52 mol of HCl, How many moles of H_2 are produced ?

[HCl is limiting reagent; H_2 formed = 0.36 mol]

15 How many moles of Lead (II) chloride will be formed from a reaction between 6.5 g of PbO and 3.2 g of HCl ? [Atomic mass of Pb = 207 U]

[**Ans.** 0.029 mole]

16 What volume of oxygen at N.T.P is needed to cause the complete combustion of 200 ml of acetylene ?Also calculate the volume of carbon dioxide formed. [**Ans.** 500 mL of O_2 & 400 mL of CO_2]

5-MARKS QUESTIONS

1 (i) A black dot used as a full stop at the end of a sentence has a mass of about one attogram. Assuming that the dot is made up of carbon, calculate the approximate number of carbon atoms present in the dot. [Hint : 1 attogram = 10^{-18}g] [**Ans.** 5.02×10^4]

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

(a) 1g Au (s) (b) 1g Na (s) (c) 1g Li (s) (d) 1g of Cl_2(g)

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

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

(ii) A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gas 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calcuate

(i) Empirical formula, (ii) molar mass of the gas, and (iii) Molecular formula. [**Ans.** (i) CH, (ii) 26 g mol^{-1}, (iii) C_2H_2]

3. (i) What is the difference between Molarity and Molality.

   (ii) The Molarity of a solution of sulphuric acid is 1.35 M. Calculate its molality. (The density of acid solution is 1.02 g cm^{-3}).

   [**Ans.**. 1.52 m]

4. (i) Define : (a) Mole fraction (b) Mass percentage.

   (ii) If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution ? [**Ans.** 0.025 L]

HOTS QUESTIONS

1 In a compound $C_xH_yO_z$, the mass % of C and H is 6:1 and the amount of oxygen present is equal to the half of the oxygen required to react completely C_xH_y. Find the empirical formula of the compound.

[**Ans.** $C_2H_4O_3$]

2 A crystalline salt when heated becomes anhydrous and loses 51.2 % of its weight. The anhydrous salt on analysis gave the following percentage composition

Mg = 20.0% , S = 26.6 % , O = 53.33 %

Calculate the molecular formula of the anhydrous salt and the crystalline salt. Molecular weight of the anhydrous salt is 120.

[**Ans.** $MgSO_4.7H_2O$]

3 An LPG cylinder weighs 14.8 kg when empty. When full, it weighs 29.0 kg and shows a pressure of 2.5 atm. In the course of use at 27°C, the weight of cylinder is reduced to 23.2 kg. Find the volume of n-butane in cubic meters used up at 27°C and 1 atm (Molecular weight of n-butane = 58). [**Ans.** 2.463 m^3]

4 2.5 g of $CaCO_3$ was placed in 50 ml of a solution of HCl.1.05 g of $CaCO_3$ was left after the reaction. Calculate:

(a) the weight of HCl per litre

(b) the Molarity of HCl [**Ans.** (a) 21.17 g, (b) 0.58 M]

UNIT TEST-I

Time allowed : 1 hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. If 30 mL of H_2 and 20 mL of O_2 react to form water, what is left at the end of the reaction ? (1)

   (a) 10 mL of H_2 (b) 5 mL of H_2

   (c) 10 mL of O_2 (d) 5 mL of O_2

2. 7.5 grams of a gas occupy 5.6 litres of volume at STP the gas is (1)

   (a) NO (b) N_2O (c) CO (d) CO_2

3. Write the relationship between empirical formula and molecular formula. (1)

4. Why molarity is preferred over molarity in expressing the concentration of solution ? (1)

5. Which has more number of atoms ? 1.0 g Na or 1.0g Mg? (1)

6. How many atoms and molecules are present in 124 g of phosphorus (P_4)? (2)

7. (a) How many significant figures are present in 0.0102. (2)

   (b) Write the number in scientific notation 1013.6.

8. A sample of drinking water was found to be severely contaminated with chloroform $CHCl_3$. The level of contamination was 15 ppm (by mass).

   (a) Express this in percent by mass.

   (b) Determine the molarity of chloroform in the water sample. (3)

9. 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 formula ? (3)

10. (a) Explain the following terms:

    (i) Gay Lussac's law (ii) Limiting reagent

    (b) 3.0 g of H_2 react with 30.0 g of O_2 yield H_2O.

    (i) Which is the limiting reagent?

    (ii) Calculate the maximum amount of H_2O that can be formed.

    (iii) Calculate the amount of reactant left unreacted. (5)

UNIT TEST-II

Time allowed : 1 hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. One mole of oxygen gas at STP is equal to -------------- (1)

   (a) 6.022×10^{23} molecules of oxygen

   (b) 6.022×10^{23} atoms of oxygen

   (c) 16 g of oxygen

   (d) 32 g of oxygen

2. 1g of M_2 CO_3 on treatment with excess HCl produces 0.01186 moles of CO_2. The molar mass of M_2CO_3 in g mol^{-1} is ? (1)

   (a) 1186 (b) 84.3 (c) 118.6 (d) 11.86

3. How many atoms are present in 16 g of ozone? (1)

In following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(a) Assertion and Reason are true and Reason is the correct explanation of Assertion

(b) Assertion and Reason are true but Reason is not the correct explanation of Assertion

(c) Assertion is true but Reason is false

(d) Both Assertion and Reason are false

4. **Assertion** : The empirical mass of ethene is half of its molecular mass. (1)

   Reason : The empirical formula represents the simplest whole number ratio of various atoms present in a compound.

5. **Assertion** : Combustion of 16 g of methane gives 18 g water. (1)

Reason : In the combustion of methane, water is one of the products.

6. If 2 litres of N_2 is mixed with 2 litres of H_2 at a constant temperature and pressure, then what will be the volume of NH_3 formed? (2)

7. Calculate the percentage of Copper in a sample of $CuCl_2$ (2)

(Atomic mass of Cu = 63.5u, Cl = 35.5u)

8. In an experiment, when HCl was reacted with $CaCO_3$ at STP, 48 Cm^3 of CO_2 was formed. Calculate the number of mole of CO_2 and number of molecules. (3)

9. In the reaction $2A + 4B \rightarrow 3C + 4D$, when 5 moles of A react with 6 moles of B, then (3)

   (i) Which is the limiting reagent

   (ii) Calculate the amount of C formed

   (iii) Calculate the amount of excess reagent left after reaction

10. (a) How many grams atoms are there in 8.0 g of S? (5)

    (b) The molarity of solution of H_2SO_4 is 1.35 M. Calculate its molality.

    (The density of solution is 1.02 g cm^{-3})

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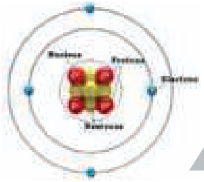


Chapter - 2

Structure of Atom

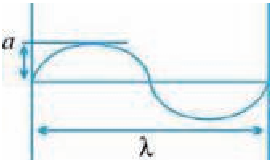
FAST TRACK : QUICK REVISION

- **Information about fundamental particles of atom**

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

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

- **Characteristics of waves :**



(a) Wavelength : The distance of one crest and one trough in a wave. Denoted by ‘λ’.

(b) Frequency : Number of waves passing through a given point in one second.

Denoted by υ.

$$\left[\begin{array}{l} \upsilon = \dfrac{1}{t} \Rightarrow \text{sec}^{-1} \text{ or Hz} \\ t = \text{Time period} \end{array}\right]$$

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

(d) Wave Number : Number of waves per unit length denoted by $\bar{\upsilon}$

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

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

$$\text{velocity of light } c = \frac{\text{Distance}}{\text{Time}} = \lambda \times \frac{1}{t} = \upsilon\lambda$$

$$\therefore \quad \upsilon = \frac{c}{\lambda}$$

- **Energywise** order for EM radiation.

$$\text{cosmic} < \gamma \text{ rays} < \text{X rays} < \text{UV} < \text{VIBGYOR} < \text{IR} < \text{Microwaves} < \text{Radiowaves}$$

$\longrightarrow$

λ (**Increases**) υ (**Decreases**) Energy (**Decreases**)

- **Photon :** A packet or particle of light energy is knows as **Photon.**
- **Planck's quantum theory :** The energy emitted or absorbed by a source is discontinuous in form of small packet of energy, called **quantum**. Quantum of light is called **photon.**

$E \propto v$

$E = hv$ (h = Planck's constant)

$E = nhv$ ($h = 6.626 \times 10^{-34}$ J sec)

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

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

$$h\upsilon - h\upsilon_0 = \frac{1}{2}mv^2$$

$h\upsilon \Rightarrow$ Energy of incident light on metal surface.

$h\upsilon_0 \Rightarrow$ Work function of metal.

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

- **de Broglie equation :** All material particles in motion also exhibit wave like properties.

$$\lambda = \frac{h}{mv} = \frac{h}{p}$$

For microscopic particles mass is very less therefore Wavelength of wave associated with it can be detected.

For macroscopic particles mass is large, λ of wave associated with it can not be detected. Hence dominant wave character.

Hence microscopic bodies have dual nature, where as macroscopic bodies have particle nature.

Heisenberg's Uncertainty Principle

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

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

Δx = uncertainty in position

Δv = uncertainty in velocity

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

(1) e^- revolving round the nucleus in circular path [stationary state; SHELL] with a definite angular momentum $\frac{nh}{2\pi}$ [Here n = no. of shell of e^-] and with definite energy

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

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

As n decreases, Energy of e^- becomes More – ve [Due to more force of attraction by protons]

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

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

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

Emission Spectrum : The emission of light energy is known as emission spectra. It corresponds to each atom depending upon which energy shell e^- is excited. It is **discontinuous** spectra as ‘λ’ of light radiations do not merge with each other like in VIBGYOR (Continous Spectra).

When e^- falls from any excited state to

$$\frac{1}{\lambda} = 1{,}09{,}678\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]Z^2 \quad R = \text{Rydberg constant} = 109678\ \text{cm}^{-1}$$

$n_i = 1, n_f = 2, 3, 4, ...$ [Lyman series] (UV)

$n_i = 2, n_f = 3, 4, 5, ...$ [Balmer series] (VIBGYOR)

$n_i = 3, n_f = 4, 5, 6$ [Paschen series] IR.

$n_i = 4, n_f = 5, 6, 7$ [Bracket series] IR.

$n_i = 5, n_f = 6, 7, 8$ [Pfund series] IR.

Quantum numbers : The numbers which **completely** define the **state** of e^- in an atom.

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

$n = 1, 2, 3, 4, 5, ...$

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

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

n	l (0 to n-1)				
1	0	$l = 0$	‘*s*’	subshell	
2	0, 1	$l = 1$	‘*p*’	subshell	
3	0, 1, 2	$l = 2$	‘*d*’	subshell	
4	0, 1, 2, 3	$l = 3$	‘*f*’	subshell	

All subshells are wave functions for locating e^-.

In the same shell energy increase $s < p < d < f$.

(3) Magnetic Quantum No. : It gives the no. of magnetic orientations an e^- can have in a subshell. That is number of orbitals in a sub-shell. $m_s = -l........0........+l = (2l + 1)$.

(4) Spin Quantum No. : An e^- is continuously spinning on its own axis.

The value of $s = \frac{1}{2}$ or $-\frac{1}{2}$

An orbital can have maximum two e^- one with clockwise and other with anticlockwise spin.

Aufbau principle

(a) Electrons are filled in increasing order of energy of sub-shell.

(b) As '$n + l$' value increases energy of e^- increases in that sub-shell.

(c) For two sub-shells with same '$n + l$' value, as 'n' value increases energy of e^- increases.

Pauli's principle

No two electrons can have same set of four quantum numbers in an atom.

Hund's rule of maximum multiplicity

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

IMPORTANT POINTS

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

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

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

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

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

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

(D) The region in an orbital where probability of finding the e^- is zero is known as **Nodal plane** (or Node).

The no. of [radial nodes] = $n - l - 1$ and Angular Nodes = l,

Total nodes = n–1.

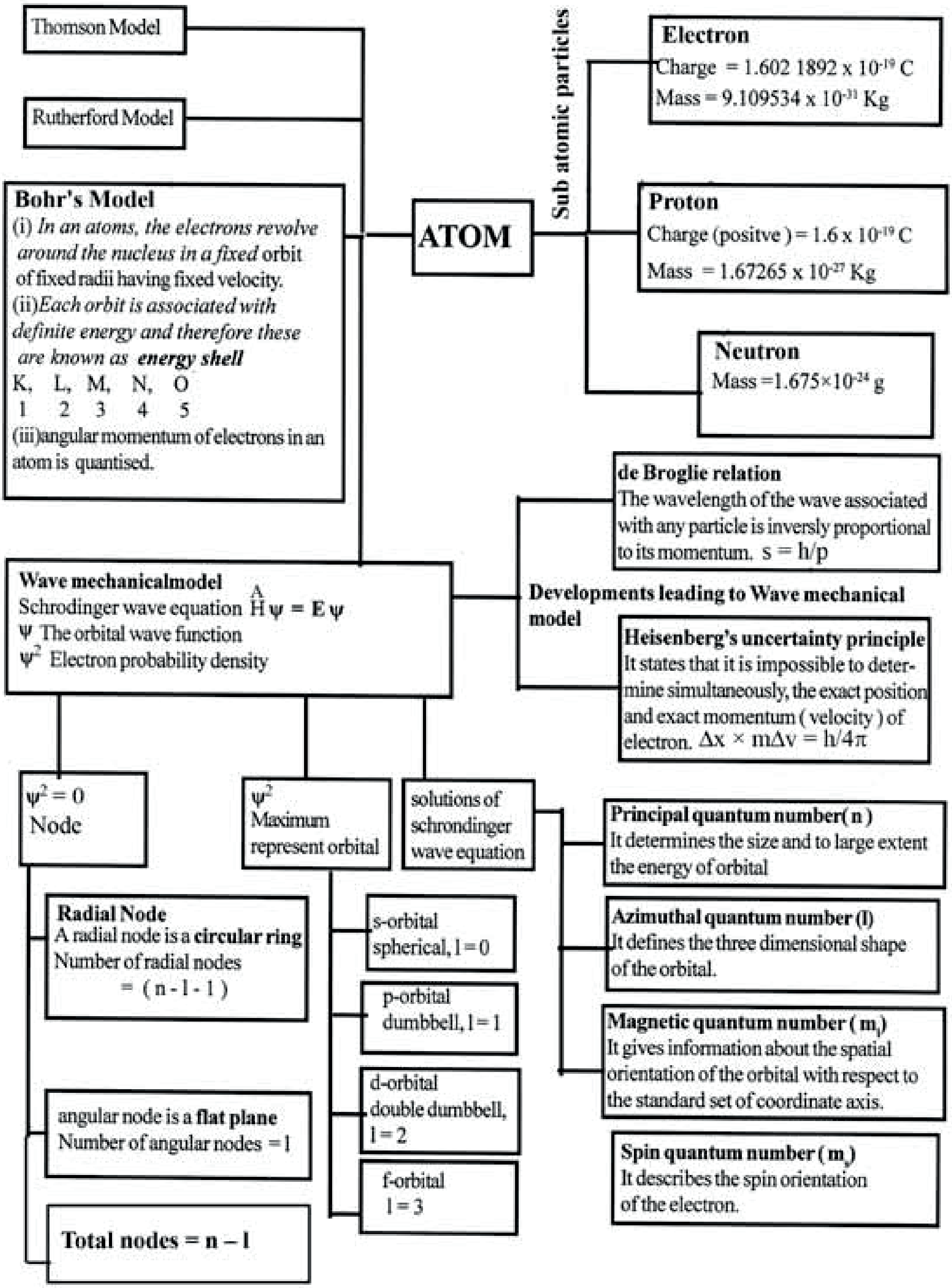
(E)

ψ(psi)	**ψ^2(psi square)**
A wave function for locating an electron	The square of wave function where the probability of finding the e^- is maximum. [Each value of ψ^2 is a region and **defines** one orbital]

(F)

(1) A definite distance from the nucleus for finding the e^- [e^- as a particle].	(1) A probability region for locating the e^- around the nucleus. It is a wave function [e^- as a wave]
(1) It has definite size and e^- in this orbit has definite energy.	(2) It does not define definite size. But only a boundary region diagram of a wave for locating the e^-.

MIND MAP - STRUCTURE OF ATOM



Thomson Model

Rutherford Model

Bohr's Model

(i) *In an atoms, the electrons revolve around the nucleus in a fixed* orbit of fixed radii having fixed velocity.

(ii)*Each orbit is associated with definite energy and therefore these are known as* ***energy shell***

K,	L,	M,	N,	O
1	2	3	4	5

(iii)angular momentum of electrons in an atom is quantised.

ATOM

Sub atomic particles

Electron

Charge = 1.602 1892 x 10^{-19} C

Mass = 9.109534 x 10^{-31} Kg

Proton

Charge (positve) = 1.6 x 10^{-19} C

Mass = 1.67265 x 10^{-27} Kg

Neutron

Mass =1.675×10^{-24} g

Wave mechanicalmodel

Schrodinger wave equation $\hat{H}\psi = E\psi$

ψ The orbital wave function

ψ^2 Electron probability density

Developments leading to Wave mechanical model

de Broglie relation

The wavelength of the wave associated with any particle is inversly proportional to its momentum. $s = h/p$

Heisenberg's uncertainty principle

It states that it is impossible to determine simultaneously, the exact position and exact momentum (velocity) of electron. $\Delta x \times m\Delta v = h/4\pi$

$\psi^2 = 0$
Node

Radial Node

A radial node is a **circular ring**

Number of radial nodes = (n - l - 1)

angular node is a **flat plane**

Number of angular nodes = l

Total nodes = n – l

ψ^2
Maximum represent orbital

s-orbital
spherical, l = 0

p-orbital
dumbbell, l = 1

d-orbital
double dumbbell, l = 2

f-orbital
l = 3

solutions of schrondinger wave equation

Principal quantum number(n)

It determines the size and to large extent the energy of orbital

Azimuthal quantum number (l)

It defines the three dimensional shape of the orbital.

Magnetic quantum number (m_l)

It gives information about the spatial orientation of the orbital with respect to the standard set of coordinate axis.

Spin quantum number (m_s)

It describes the spin orientation of the electron.

CASE BASED QUESTIONS

1. **Read the passage given below and answer the following questions:**

The capacities of shells with a given principal quantum number are fixed by (1) the rules governing the permitted values of the quantum numbers and (2) the Pauli Exclusion Principle. The permitted values of the quantum numbers are :

Principal quantum number	n	1 to ∞
Azimuthal quantum number	l	0 to $n-1$ (n values)
Magnetic quantum number	m_l	$-l$ to $+l$, ($2l+1$ values)
Spin quantum number	m_s	$-½$ or $+½$ (2 values)

The Pauli Exclusion Principle states that no two electrons in the same atom may have the same values of all four quantum numbers. It follows that, for a given value of n, there are $2n^2$ different sets of values for the quantum numbers, because l may have the values O, 1, . . ., $n-1$, and for each value of l there are $2l+1$ values of m_l and for each set of values of l and m there are just two choices for m_s.

(Reference : Thomas H. Hazlehurst, J. Chem. Educ. 1941, 18, 12, 580 Publication Date: December 1, 1941, Journal of American Chemical Society).

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) Based on the concept of quantum numbers mentioned in the study, which of the following value sets of quantum numbers are not possible?

(a) n=1, l=0, m_l= 0, m_s= -1/2

(b) n=5, l=3, m_l= -4, m_s= +1/2

(c) n=3, l=1, m_l= -1, m_s= -1/2

(d) n=6, l=1, m_l= 0, m_s= +1/2

(ii) What will be the maximum possible number of electrons having m_s = -1/2 for n=5?

(a) 50 (b) 25

(c) 32 (d) 72

(iii) Which of the following quantum numbers can distinguish between two electrons present in the same orbital?

(a) Azimuthal quantum number

(b) Principal quantum number

(c) Magnetic quantum number

(d) Spin quantum number

(iv) Maximum number of electrons having n = 3 and l = 1 is-

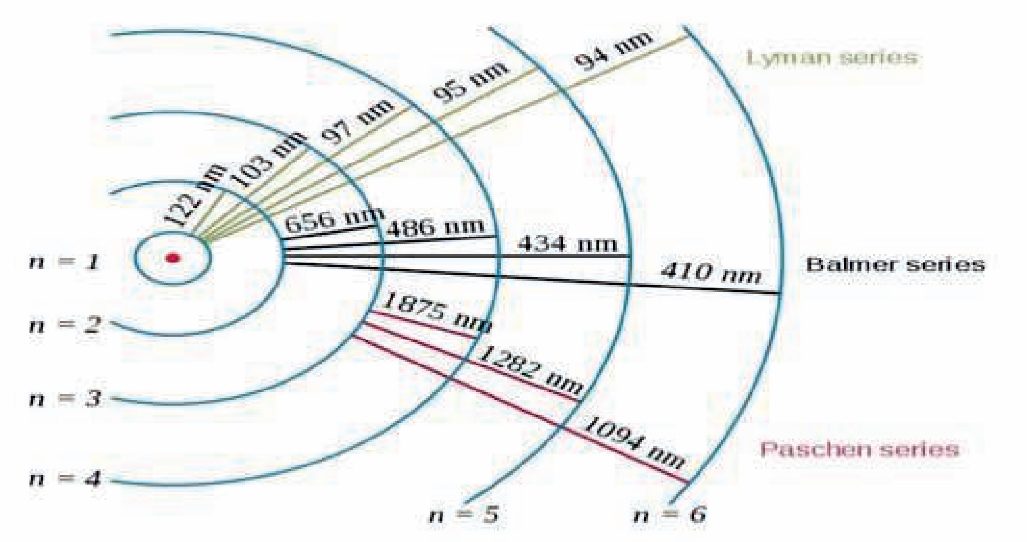
(a) 14 (b) 6

(c) 10 (d) 2

ANS: I-B, II-B, III-D, IV-B

2. Read the passage given below and answer the following questions:

A hydrogen atom consists of an electron orbiting its nucleus. The electromagnetic force between the electron and the nuclear proton leads to a set of quantum states for the electron, each with its own energy. These states were visualized by the Bohr model of the hydrogen atom as being distinct orbits around the nucleus. Each energy state, or orbit, is designated by an integer, *n* as shown in the figure. The Bohr model was later replaced by quantum mechanics in which the electron occupies an atomic orbital rather than an orbit, but the allowed energy levels of the hydrogen atom remained the same as in the earlier theory.



Spectral emission occurs when an electron transitions, or jumps, from a higher energy state to a lower energy state. To distinguish the two states, the lower energy state is commonly designated as n', and the higher energy state is designated as n. The energy of an emitted photon corresponds to the energy difference between the two states. Because the energy of each state is fixed, the energy difference between them is fixed, and the transition will always produce a photon with the same energy.

(Reference : Andrew, A. V. (2006). "2. Schrödinger equation". Atomic spectroscopy Introduction of theory to Hyperfine Structure p. 274 ISBN 978-0-387-255736. https://en.wikipedia.org/wiki/Hydrogen_spectral_series).

In these questions (Q. No. (i) to (iv) , a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices :

(a) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(c) Assertion is correct statement but reason is wrong statement.

(d) Assertion is wrong statement but reason is correct statement

i. ASSERTION: The energy states of H-atom are independent of azimuthal quantum number.

REASON: H-atom does not have any inter-electronic repulsion as it has only 1 e^-.

ii. ASSERTION: A spectral line will be seen for a $2p_x - 2p_y$ transition.

REASON: Energy is released when electron drops to lower stationary state.

iii. ASSERTION: For Balmer series of hydrogen spectrum, the value $n_1 = 2$ and $n_2 = 3, 4, 5...$.

REASON: The value of n_2 for a line in Balmer series of hydrogen spectrum having the highest wavelength is 6.

iv. ASSERTION: Electromagnetic radiations of fixed wavelengths are absorbed by the H-atom.

REASON: Radiations corresponding to the energy difference between the two stationary states are absorbed.

ANS:- I-A, II-D, III-C, IV-A

1. Packet of energy is called
   (a) Electron (b) Photon
   (c) Position (d) Proton
2. Orbital which is not possible
   (a) 2p (b) 3d
   (c) 3s (d) 3f
3. the magnetic quantum number of an atom is related to the
   (a) size of the orbital (b) spin angular momentum
   (c) orbital angular momentum (d) orientation of the orbital in space
4. The principal quantum number of an atom is related to the
   (a) size of the orbital (b) spin angular momentum
   (c) orbital angular momentum (d) orientation of the orbital in space
5. The designation of an orbital with n = 4 and 1 = 3
   (a) 4s (b) 4p
   (c) 4d (d) 4f
6. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 in the He^+ spectrum?
   (a) n = 4 to n =1 (b) n = 3 to n = 2
   (c) n = 3 to n = 1 (d) n = 2 to n = 1
7. The wave number of first line of Balmer series of hydrogen in 15200 cm-1. The wave number of the first Balmer line of Li^{2+} ion is
   (a) 15200 cm^{-1} (b) 60800 cm^{-1}
   (c) 76000 cm^{-1} (d) 136,800 cm^{-1}
8. An electron is moving in Bohr's orbit. Its de Broglie wavelength is λ. What is the circumference of the Third orbit?
   (a) $2/\lambda$ (b) 2λ
   (c) 3λ (d) $3/\lambda$

9. Which of the following statements in relation to the hydrogen atom is correct?

(a) 3s-orbital is lower in energy than 3p-orbital

(b) 3p-orbital is lower in energy than 3-d-orbital

(c) 3s and 3p orbitals all have the same energy.

(d) 3s, 3p and 3d orbitals all have the same energy.

10. For principle quantum number, n = 4, the total number of orbitals having 1 = 3 is

(a) 3 (b) 7

(c) 5 (d) 9

11. The number of d-electrons retained in Fe^{2+} (At. no. of Fe = 26) ion is

(a) 3 (b) 4

(c) 5 (d) 6

12. Pauli exclusion principle helps to calculate the maximum number of electrons that can be accommodated in any

(a) orbital (b) subsell

(c) shell (d) All of these

Ans. 1. (b), 2. (d), 3. (d), 4. (a), 5. (d), 6. (d), 7. (d), 8. (c), 9. (d), 10. (b), 11. (d), 12. (a)

FILL IN THE BLANK

1. Bohr's theory is based on ________ of radiation.

2. The angular momentum of the electron in the 4th energy shell in the hydrogen atom is ________.

3. Lines of Balmer series appear in ________ region.

4. The maximum number of electrons in Fe^{3+} (At. No. 26) is ________.

5. Li^{2+} and He^{+} ions have spectrum similar to ________ atom.

6. Bohr's atomic theory is not able to explain the atomic spectra of atoms containing ________ electron.

7. An electron in the first shell will have ________ stability and ________ energy than an electron in the third shell.

8. The space or three-dimensional region round the nucleus where there is maximum probability of finding an electron of specific energy is called an_____

9. According to _____ no two electrons in an atom will have all the four quantum numbers _____

10. When there are two electrons in the same orbital they have _____ spins.

11. The s-subhells have _____ shape and the p-subshells have_____

12. The maximum number of orbital on a subshell is equal to _______ where l= _______

Ans. 1. Planck's theory

2. $\frac{2h}{\pi}$

3. Visible

4. 23

5. H-atom

6. more than 1

7. Larger, lower

8. orbital

9. Pauli exclusion principle; similar

10. Opposite

11 Spherical, dumb bell shape.

12. 2l + 1; azimuthal quantum numbers

TRUE AND FALSE TYPE QUESTIONS

Write true or false for the following statements

1. Bohr's theory cannot explain the spectra of multi-electron atoms.
2. Bohr's theory based on the Planck's quantum theory.
3. Size of orbital is determined by principal quantum number.
4. Fe^{2+} ion has more number of unpaired electrons than Fe^{3+}.
5. The outer electronic configuration of chromium atom is $3d^4 4s^2$.
6. The designation of an orbital n=4 and l=0 is 4s.
7. All photons of light have same energy.
8. Fe^{3+} has $3d^5$ configuration.

9. The number of subshells is always equal to the order of the orbit.

10. Two electrons in the same orbital has antiparallel spin.

11. The second orbit in He^+ ion has radius as the first orbit in hydrogen atom.

12. Heisenberg principle is applicable to microscopic particles.

13. 3s orbital has 2 radial nodes.

Ans. 1. (T) 2. (T) 3. (T) 4. (F) 5. (F) 6. (T) 7. (F)
8. (T) 9. (F) 10. (T) 11. (F) 12. (T), 13. (T)

MATCH THE COLUMNS

1. Match the following

	List-I		List-II
a.	Lyman series	p.	Visible region
b.	Balmer series	q.	Infrared region
c.	Paschen series	r.	Absorption spectrum
d.	Brackett series	s.	Ultraviolet region

2. Match the following

	List-I		List-II		List-III
a.	Principal quantum number	p.	Spin of electrons	i.	$-l$ to $+l$
b.	Azimuthal quantum number	q.	Size of orbital	ii.	0 to ∞
c.	Magnetic quantum number	r.	Orientation of the orbital	iii.	$\pm \frac{1}{2}$
d.	Spin quantum number	s.	Shape of the orbital	iv.	0 to $(n-1)$

3. Match the following

	List-I		List-II		List-III
a.	2s	p.	DOUGHNUT shape	i.	along z-axis
b.	$2p_x$	q.	Spherical	ii.	In between x & y-axis
c.	$3d_{xy}$	r.	Dumb bell	iii.	non-directional
d.	$3d_{z^2}$	s.	Double dumb bell	iv.	along x-axis

4. Match the following

	List-I		**List-II**
a.	2s	p.	Two nodal planes
b.	ψ^2	q.	One radial node
c.	Heisenberg's uncertainty principle	r.	Electron probability density
d.	$3d_{yz}$	s.	Microscopic particles

Ans.: 1. a. (s), b. (p), c. (q), d. (q)

2. a. (q). (ii), b. (s). (iv), c. (r). (i), d. (p). (iii)

3. a. (q). (iii), b. (r). (iv), c. (s). (ii), d. (p).(i)

4. a. (q), b. (r), c. (s), d. (p)

ASSERTION AND REASON TYPE QUESTIONS

Directions: (Questions 1 to 10)

A. Both Assertion & Reason are true and the reason is the correct explanation of the assertion.

B. Both Assertion & Reason are true but the reason is not the correct explanation of the assertion.

C. Assertion is true statement but Reason is false.

D. Assertion is false but Reason is true.

1. **Assertion :** Number of orbitals in 3rd shell is 9.

   Reason : Number of orbitals for a particular value of n = n^2.

2. **Assertion :** Two nodal planes are present in $3d_{xy}$.

   Reason : Number of nodal planes = l

3. **Assertion :** The energy of an electron is largely determined by its principal quantum number.

   Reason : The principal quantum number is a measure of the most probable distance of finding the electrons around the nucleus.

4. **Assertion :** An orbital cannot have more than two electrons, moreover, if an orbital has two electrons they must have opposite spins.

   Reason : No two electrons in an atom can have same set of all the four quantum numbers.

5. Assertion: Black body is an ideal body that emits and absorbs radiations of all frequencies.

   Reason: The frequency of radiation emitted by a body goes from a lower frequency to higher frequency with an increase in temperature.

6. Assertion: 2p orbitals do not have any radial nodes.

   Reason: The number of radial nodes in p-orbitals is given by (n-2) where n is the principal quantum number.

7. Assertion: The opposite lobes of a p-orbital have opposite sign whereas opposite lobes of d-orbital have the same sign.

   Reason: The opposite lobes of a p-orbital have opposite charge whereas the opposite lobes of d-orbital have the same charge.

8. Assertion: Electronic configurations of Cr^{3+} (containing 21 electrons) is same as that of Sc(Z=21) i.e., isoelectronic species have the same electronic configuration.

   Reason: Orbitals of atoms are filled in order of increasing energy following aufbau principle.

9. Assertion: Hydrogen has one electron in its orbit but it produces several spectral lines.

   Reason: There are many excited energy levels available.

10. Assertion: The free gaseous Cr atom has six unpaired electrons.

    Reason: Half-filled d-orbitals have greater stability.

Ans. 1. A 2. A 3. A 4. A 5. B 6. A 7. C 8. D 9. A 10. A

ONE WORD ANSWER TYPE QUESTIONS

1. Write the name of the theory which explain the wave nature of light.

2 Write the name of the theory which explain the Black body radiations and photo electric effect

3 If the length of the crest of a wave is 4 pm. Write the wavelength of this wave. [Ans.8 pm]

4. A radiation emitted from a hot iron is photon or quantum ?

5. Out of the d orbitals which does not have four lobes ?

6. What is the lowest value of n that allows g orbitals to exist ?

7. Which quantum number is not obtained from solution of Schrödinger wave equation ?

8. Which of the following orbitals are possible ?

1p, 2s, 2p and 3f

9. Write the name of non-directional subshell.

10. Write the name of quantum number which determines the orientation of orbitals ?

11. Write the name of quantum number which determines the shape of orbitals.

12. How many orbitals are present in 'g' subshell ?

1-MARK QUESTIONS

1 Write the relation between frequency and wave number.

2 Cs shows maximum photoelectric effect, why ?

3 Distinguish between a photon and a quantum.

4 The line spectrum of an element is known as fingerprints of its atom. Comment.

5 What is the value of the Bohr's radius for the third orbit of hydrogen atom?

6 What type of metals are used in photoelectric cell ? Give one example. [**Ans.** With large size, less work function.]

7 Which series of lines of the hydrogen spectrum lie in the visible region'?

8. What is uncertain in uncertainty principle ?

9. Can a moving cricket ball have a wave character ? Justify your answer.

10. Heisenberg uncertainty principle has no significance in our everyday life. Explain.

11. Why uncertainty in position is more when uncertainty in velocity is less for an electron ?

12. What are the four quantum numbers of 19th electron of copper ?

(Given : Atomic number of copper = 29)

13. How many electrons will be present in the sub-shells having ms, value of – 1/2 for n = 4 ?

14. Write the electronic configuration of Ni^{3+}. (At. No. of Ni = 28)

15. How many radial and angular nodes are present in 2p orbital.

16. Mention the physical significance of Ψ and Ψ^2.

[**Ans.** Radial nodes = 0, Angular nodes = 1]

2-MARKS QUESTIONS

Q. 1. Define black body and black body radiations.

Q. 2. Give the essential postulates of Bohr's model of an atom. How did it explain ?

(i) the stability of the atom ?

(ii) origin of the spectral lines in H-atom ?

Q. 3. What is quantisation ? How quantisation of energy was introduced in Bohr's model ?

Q. 4. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum?

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

Q. 5. What transition of Li^{2+} spectrum will have the same wavelength as that of the second line of Balmer series in He^+ spectrum ?

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

Q. 6. Calculate the energy required for the process

$$He^+(g) \longrightarrow He^{2+}(g) + e^-$$

The ionization energy for the H atom in the ground state is 2.18×10^{-18} J $atom^{-1}$] [**Ans.** 8.72×10^{-18} J]

Q. 7. Calculate the wave number for the longest wavelength transition in the Balmer series of atomic hydrogen. [**Ans.** 1.523×10^{6} m^{-1}]

Q. 8. To which orbit the electron in H atom will jump on absorbing 12.1 eV energy ? [**Ans.** 3rd orbit]

Q. 9. Calculate the energy associated with the first orbit of He^{+}. What is the radius of this orbit? [**Ans.** – 54.38 eV, 0.2645 Å]

Q. 10. What is the distance of separation between 3rd and 4th orbit of H-atom ? [**Ans.** 3.703 Å]

Q. 11. The energy of electron in the first Bohr's orbit is – 13.6 eV. Calculate the energy of electron in the first excited state. [**Ans.** – 3.4 eV]

Q. 12. Calculate the number of protons emitted in 10 hours by a 60 W sodium lamp emitting radiations of wavelength 6000 Å.

Q. 13. Which one has a higher energy, a photon of violet light with wavelength 4000 Å or a photon of red light with wavelength 7000 Å ?

[Given. $h = 6.62 \times 10^{-34}$ *J sec.]*

Q. 14. A 100 watt bulb emits monochromatic light of wavelength 400 nm. Calculate the number of photons emitted per second by the bulb.

[**Ans.** 2.012×10^{20} s^{-1}]

Q. 15. What are the maximum number of emission lines when the excited electron of a H atom in $n = 4$ drops to the ground state ? [**Ans.** 6]

Q. 16. Which has more energy, light radiation of wavelength 400 pm or light radiation of frequency 10^{15} Hz ?

Q. 17. Find the energy of electron in 4th shell of Li^{2+} ion.

Q. 18. What is the wave number of an electron with shortest wavelength radiation in Lyman spectrum of He^{+} ion?

Q. 19. Write short note on :

(a) Continuous and discontinuous spectrum.

(b) Absorbtion and emission spectrum.

Q. 20. Calculate the mass of the photon with wavelength of 3.6 Å.

[**Ans.** 6.135×10^{-33} kg]

Q. 21. Calculate the mass of the photon with wavelength of 5 pm.

Q. 22. On the basis of uncertainty principle show that an electron cannot exist with in atomic nucleus. *(Given : Nuclear radius = 10^{-15} m)*

[***Hint :*** *Taking 10^{-15} m as Δx, the Δv comes much higher than the velocity of light and hence is not possible*]

Q. 23. Explain why the uncertainty principle is significant only from the motion of subatomic particles and is negligible for macroscopic particles?

Q. 24. List two differences between orbit and orbital .

Q. 25. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit

Q. 26. Comment on "Bohr's model is against the Heisenberg uncertainty principle".

Q. 27. What are the similarities and difference in 2*s* and $2p_x$ orbitals and 1s and 2*s* orbitals ?

Q. 28. Draw shape of $d_{x^2-y^2}$ orbital.

Q. 29. On the basis of Pauli's exclusion principle show that the maximum number of electrons in the M -shell ($n = 3$) of any individual atom is 18.

Q. 30. Designate each subshell with $n = 4$.

Q. 31. List the possible values for all the quantum numbers for the following subshell.

(a) 2*p* (b) 4*f*

Q. 32. Write down the electronic configuration of Fe^{3+} and Ni^{2+}. How many unpaired electrons are present? (Given Atomic number, Fe = 26, Ni = 28).

Q. 33. Out of principal, angular, magnetic and spin quantum number, which quantum number determines the ?

(a) Shape of the orbital
(b) Number of orbitals in an orbit
(c) Size of the orbital
(d) Spin orientation of the electron.

Q. 34. What is the Hund's rule of maximum multiplicity ? Explain with suitable example.

Q. 35. Explain why :

(a) The three electrons present in 2p subshell of nitrogen remain unpaired.

(b) Cr has configuration $3d^5\,4s^1$ and not $3d^4\,4s^2$.

Q. 36. (a) What is difference between 'l' and 'L'?

(b) Nitrogen has 7 proton, 7 electron and 7 neutrons. Calculate the number of electron, protons and neutrons in N^{3-} ion.

Q. 37. Which one is having higher energy?

(a) Last electron of Cl^- or last electron of O^{2-}.

(b) $n = 4, l = 3$ or $n = 5, l = 2$.

3-MARKS QUESTIONS

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

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

[**Ans.**– 1.36×10^{-19} J atom^{-1} .4.761Å]

Q. 2. A bulb emits light of wave length 4500Å. The bulb is rated as 150 watt and 8% of the energy is emitted as light. How many photons are emitted by the bulb per second ? [**Ans.** $n = 27.2 \times 10^{18}$]

Q. 3. When light with a wavelength of 400 nm falls on the surface of sodium, electrons with a kinetic energy of 1.05×10^5 J mol^{-1} are emitted.

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

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

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

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

Q. 5. Calculate the number of waves made by Bohr electron in one complete revolution in its third orbit. [**Ans.** 3]

Q. 6. What should be the ratio of velocities of CH_4 and O_2 molecules so that they are associated with de Broglie waves of equal wavelength ? [**Ans.** 2]

Q. 7. Calculate the wavelength of an electron that has been accelerated in a particle accelerator through a potential difference of 1 kv.

[Given 1eV = 1.6×10^{-19} J] [**Ans.** 3.87×10^{-11} m]

Q. 8. (i) Discuss the similarities and differences between a 1s and 2s orbital.

(ii) Draw the shape of d_{z^2} .

Q. 9. Calculate the wavelength of a tennis ball of mass 60 gm moving with a velocity of 10 m per second. *($h = 6.626 \times 10^{-34}$ kg m^2 s^{-1})*

[**Ans.** 10^{-33}metre]

Q. 10. Calculate the wavelength of 1000 kg rocket moving with a velocity of 3000 km/hr. *($h = 6.626 \times 10^{-34}$ kg m^2 s^{-1})*

[**Ans.** 7.9512×10^{-40} m]

Q. 11. Calculate the uncertain it in the velocity of a cricket ball of mass 150 g, if uncertainity in its position is of the order of 1 Å.

[**Ans.** 3.5×10^{-24} m s^1]

Q. 12. (a) What is de-Broglie wavelength for an electron moving with velocity of light?

(b) What is the angular momentum of electron in 5th shell?

Q. 13. Two particles A and B have wavelength $\lambda_A = 5 \times 10^{-10}$ m and $\lambda_B = 10 \times 10^{-10}$m. Find their frequency, wave number and energies. Which has more penetrating power and why ?

Q. 14. (a) Which has max. uncertainty regarding position and why ?

Electron, proton and neutron.

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

Q. 15. If an electron in He^+ has angular momentum of $5h/2\pi$. Find its energy and wavelength associated with it. Find the kinetic energy of this electron.

Q. 16 .(i) An atomic orbital has $n = 2$. What are the possible values of l and m_l ?

(ii) List the quantum numbers (m_1 and l) of electrons for 3*d* orbital.

(iii) Which of the following orbitals are possible ?

2*d*, 1*s*, 2*p* and 3*f*.

Q. 17. (a) Write the maximum number of electron in a subshell with $l = 3$ and $n = 4$.

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

$n = 3, l = 1$ and $m_l = -1$

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

Q. 18. (i) Write the electronic configurations of the following ions :

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

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

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

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

Q. 19. Calculate:

(a) Total number of spherical nodes in 3*p* orbital.

(b) Total number of nodal planes in 3*p* orbital.

(c) Nodal planes in 3*d* orbital.

5-MARKS QUESTIONS

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

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

[**Ans.** $\nu_o = 4.41 \times 10^{14}\ s^{-1}$ $W_o = 2.92 \times 10^{-19}$ J]

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

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

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

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

(i) Threshold frequency (ii) Work function.

(b) The work function for Cs atom is 1 .9 eV. Find threshold wavelength (λ_0)and threshold frequency (ν_0) of this light radiation. If Cs metal is irradiated with a radiation of wavelength 500 nm find kinetic energy and velocity of emitted electron.

Q. 4.(a) State de Broglie equation. Write its significance.

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

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

Q. 5.(a) State Heisenbergs uncertainty principle. Give its mathematical expression. Also give its significance.

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

[**Ans.** 9.55×10^{-10}m]

Q. 6.(a) Cricket ball, a tennis ball and a proton which has more uncertainity in velocity and which follows Heisenberg uncertainity principle maximum.

(b) What is the similarity in de-Broglie and Heisenberg principle? Which is different from Bohr theory for structure of atom?

(c) Why energy in a given subshell is negative?

Q. 7.(a) Write short notes on:

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

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

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

Q. 8.(a) Draw the shapes of the following orbitals.

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

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

(c) Using *s*, *p*, *d*, *f* notations, describe the orbital with the following quantum numbers:-

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

Q .9. Explain the following :

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

(ii) Maximum number of electron with – 1/2 spin for $n = 3$ is 6,9,12 or none.

(iii) Maximum number of electron can be present for $n + l = 4$.

(iv) 3*f* subshell is not possible.

(v) Maximum number of electrons in a subshell is :

$(2l + 1)$ or $(4l + 1)$ or n^2

Q. 10.(a) A neutral atom has 2K, 8L and 15 M electrons. Find the total numbers of electrons in *s*, *p*, *d* and *f* subshell.

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

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

(Given Atomic number : Al=13, Cr = 24, Co = 27 & Mn = 25)

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

Q. 11. Answer the following :

(a) $n + l$ value for 14th electron in an atom.

(b) Increasing order of filling electron in 4*f*, 5*p* and 6*d* subshells.

(c) '*m*' and '*l*' value for last electron of Mg atom.

(Given atomic number of Mg is 12)

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

(Given Atomic number of Ga is 31)

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

UNIT TEST-I

Time allowed : 1 Hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. Designation for an orbital with $n = 4$ and $l = 3$ is (1)

   (a) 4s (b) 4p (c) 4d (d) 4f

2. Maximum number of unpaired electrons in chromium is (1)
   (Given: Atomic number of Cr = 24)

   (a) 4 (b) 5 (c) 6 (d) 7

3. Which series of lines of the hydrogen spectrum lie in the visible region'? (1)

4. Why de-Broglie's wavelength is not significant for macroscopic objects.(1)

5. Which of the following is not possible ?

   (a) 2p (b) 3d (c) 3f (d) 4p (1)

6. Write two difference between orbit and orbital. (2)

7. Calculate the wave number for the longest wavelength transition in the paschen series of atomic hydrogen. (2)

8. (a) How many orbitals are associated with $n = 4$? (3)

   (b) How many electrons will be present in the sub-shells having ms value of $-1/2$ for $n = 3$?

   (c) Draw the shape of d_{z^2}.

9. Calculate the uncertainty in the position of a dust particle with mass equal to 1 mg if the uncertainty in its velocity is 5.5×10^{-20} ms^{-1}. (3)

10. (i) The energy associated with the first orbit in the hydrogen atom is -2.18×10^{-18} J atom^{-1}. What is the energy associated with the fifth orbit?

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

    (iii) Calculate the radial and angular nodes in 2p orbital.

    (iv) Define the black body and black body radiations. (5)

UNIT TEST-II

Time allowed : 1 Hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

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1. The de-Broglie wavelength associated with a ball of mass 1 kg having kinetic energy 0.5 J is (1)

   (a) 6.626×10^{-34} m (b) 13.20×10^{-34} m

   (c) 10.38×10^{-21} m (d) 6.626×10^{-34} Å

2. The radius of which of the following orbit is same as that of first orbit of hydrogen atom? (1)

   (a) He^+ (n = 2) (b) Li^{2+} (n = 2) (c) Li^{2+} (n = 3) (d) Be^{3+} (n = 2)

3. Which series of hydrogen spectrum lies in the UV region'? (1)

In following questions a statement of Assertion followed by a statement of Reason is given. Choose the correct answer out of the following choices.

(a) Assertion and Reason are true and Reason is the correct explanation of Assertion.

(b) Assertion and Reason are true but Reason is not the correct explanation of Assertion.

(c) Assertion is true but Reason is false.

(d) Both Assertion and Reason are false.

4. **Assertion :** It is impossible to determine the exact position and exact momentum of an electron simultaneously.

   Reason : The path of an electron in an atom is clearly defined. (1)

5. **Assertion :** All isotopes of a given element show the same type of chemical behaviour.

   Reason : The chemical properties of an atom are controlled by the number of electrons in the atom. (1)

6. Calculate the number of angular nodes and radial nodes in 3p orbital. (2)

7. Calculate the mass of photon with wavelength 3.6A°. (2)

8. What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition n = 4 to n = 2 of He^+ spectrum? (3)

9. (a) The energy associated with Bohr's first orbit is -2.18×10^{-18} J $atom^{-1}$. What is the energy associated with fifth orbit?

   (b) The work function for Caesium atom is 1.9eV. Calculate the threshold wavelength.

   [Given : lev = 1.6×10^{-19} J]

   (c) How many sub-shells are associated with n = 4? (1×3=3)

10. (i) How many electrons will present in sub-shell having spin quantum number value of –½ for n = 4?

    (ii) Which of the following transition will have minimum wavelength and why?

    $n_4 \rightarrow n_1$, $n_4 \rightarrow n_2$, $n_2 \rightarrow n_1$

    (iii) Give the number of radial nodes for 3s and 2p orbitals. (5)

******

Chapter - 3

Classification of Elements and Periodicity in Properties

FAST TRACK : QUICK REVISION

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

1. **Mendeleev's periodic law**
   "The physical and chemical properties of elements are periodic functions of their atomic weight."
2. It was modified to **Modern Periodic law :**
   "The physical and chemical properties of elements are periodic functions of their atomic numbers."
   It is the long form of periodic table :
   7 Horizontal rows are called Periods and 18 Vertical columns are called Group
   Group-1 are called **Alkali metals** Group-2 are called **Alkaline earth metals**.
   Group-15 are called **Pnicogens** Group-16 are called **Chalcogens**
   Group-17 are called **Halogens** Group-18 are called **Noble gases**
3. 1st period – 2 elements 2nd and 3rd period – 8 elements
   4th and 5th period – 18 elements 6th period – 32 elements
   7th period – Incomplete (32 elements)
4. Groups
   1 and 2 – '*s*' block elements last electron entered in '*s*' subshell [s^1, s^2]
   3 to 12 – '*d*' block elements last electrons entered in '*d*' subshell [d^1 to d^{10}].
   13 to 18 – '*p*' block elements last electrons enter in '*p*' subshell [p^1 to p^6].
   Two *f*-block series lanthanoids and actinoids are placed in the bottom of periodic table.

5. (A) In '*s*' and '*p*' block elements the last electrons enters in outer most shell. In '*d*' block elements the last electron enters in the penultimate shell ($n-1$). '*f*' block elements last electron enter the antepenultimate 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]

6. **General outer electronic configuration**

's' block : ns^1, ns^2 [Group 1 to 2]

'p' block : ns^2np^1 to $ns^2\,np^6$ Group 13 to 18

'd' block : ns^{0-2} $(n-1)\,d^{1 \text{ to } 10}$ Group 3 to 12

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

7. **General periodic trends in properties of elements**

- 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.

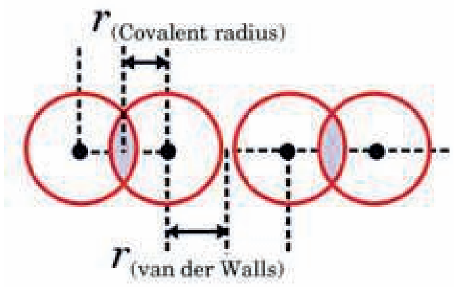
(C) Noble gases have large radius than **group 17** due to complete filling of electron in outer shell electron-electron repulsion mildly increases.

- COVALENT RADIUS

It is half of the distance between the centre of nuclei of two adjacent similar atoms which are bonded to each other by single covalent bond.

- **van der Waal's Radius**

van der Waal's radius is defined as one-half the distance between the centres of nuclei of two nearest like atoms belonging to two adjacent molecules of the element in the solid state.



- **METALLIC RADIUS**

  Half of the distance between the centres of the nuclei of two adjacent atoms in the metallic crystal. A comparison of the three atomic radii show that van der Waal's radius is maximum while the covalent radius has the least value.

  van der Waal's radius > Metallic radius > Covalent radius

- **IONIC RADIUS**

  (A) Cation radius < Atomic radius – due to more no. of protons than number of electron coloumbic force increases, size decreases.

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

  (B) Anion radius > Atomic radius – Due to more number of electron than number of protons

  $$[N^{3-} > O^{2-} > F^{-}]$$

  Electron-Electron repulsion increase, coloumbic force of attraction decreases.

  (C) For Isoelectronic species – More is the charge of cation lesser the size.

  More is the charge of anion, more is the size.

  (D) Order of size – $O^{2-} > F^{-} > Na > Na^{+} > Mg^{2+}$

8. (A) Ionisation enthalpy :

The minimum amount of energy which is required to remove the most loosely bound electron from an isolated atom in the gaseous state is called Ionisation enthalpy.

$$M(g) + \text{Energy} \longrightarrow M^{+} + e^{-}$$

$$IE_3 > IE_2 > IE_1$$

(B) Variation of I.E along a period:

Ionisation enthalpy increase along the period because atomic radii decrease and nuclear charge increase along the period.

I ionisation enthalpy Li < B < Be < C < O < N < F < Ne

II ionisation enthalpy Be < C < B < N < F < O < Ne

(C) Variation down the group:

Ionisation enthalpy decreases down the group because atomic radius increase down the group.

Metallic behaviour : Decrease from left to right due to increase in ionisation enthalpy.

Non metallic behaviour : Increase from left to right due to decrease of electron in outershell and added electron goes towards nucleus.
in atomic size and increase in effective nuclear change.

9. Screening effect or shielding effect:-

It is the decrease in the force of attraction between nucleus and outermost electron due to presence of inner shell electrons. As a result, the outer most electrons does not feel full charge of the nucleus. The actual charge felt by an electron is called effective Nuclear charge.

Shielding effect is in the following order **s > p > d > f**

d & f subshell show weak sheilding effect because their orbital size are large and are more diffused.

10. Isoelectronic species:

Ions of different elements which have the same number of electrons but different no. of protons are called isoelectronic ions.

	Na^+	Mg^{2+}	Al^{3+}	N^{3-}	O^{2-}	F^-
No. of Protons	11	12	13	7	8	9
No. of electrons	10	10	10	10	10	10

Ionic Radii $Al^{3+} < Mg^{2+} < Na^+ < F^- < O^{2-} < N^{3-}$

11. Electron gain enthalpy: ΔegH

The enthalpy change when an extra electron is added to neutral gaseous atom to form anion.

$$E(g) + e^- \longrightarrow E^-(g)$$

- **Trends : From left to right –** Increase due to decrease in size, more attraction of added electron by nucleus.
- **From top to bottom—**Decreases as the added electron is away from nucleus due to increase in size.
- ***Cl* has more negative electron gain enthalpy than fluorine –** Due to small size of fluorine extra added electron has more inter electronic repulsion than chlorine which has large size.
- Similarly Phosphorus and Sulphur have negative electron gain enthalpy than nitrogen and oxygen respectively.
- **Maximum electron gain enthalpy** – Chlorine (in periodic table)

- **Electron gain enthalpy –**
  Halogen > Oxygen > Nitrogen > Metal of group 1 and 13 and non metal of group 14 > metal of group 2.
- 2nd electron gain enthalpy is always positive.

12. Electro negativity:

The tendency of an atom to attract the shared pair of electron towards itself in a bonded state.

- Fluorine is the most electronegative element in the periodic table.
- Cesium is the least electronegative element in the periodic table.
- Electro-negativity decreases down the group and increases along the period

Difference between electron gain enthalpy and Electronegativity.

Electron gain enthalpy is the energy, but electronegativity is not the energy, it is only the tendency of an atom in a molecule to attract the shared pair of electrons. Three highest electronegative atoms F > O > N.

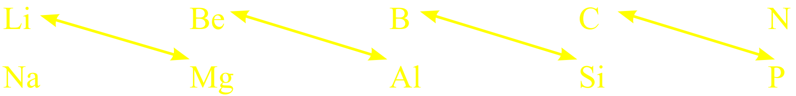
Maximum electronegative Assign to F.

* Lightest element : **Hydrogen**
* Lightest metal : **Lithium**
* Heaviest metal (highest density) : **Osmium**
* Most reactive metal : **Caesium**
* Most reactive nonmetal : **Fluorine**
* Most malleable metal : **Gold**
* Electrically best conductor : **Silver**
* Metals which are relatively volatile : **Zn, Cd, Hg**
* Strongest reducing agent in aqueous solution : **Lithium**
* Strongest oxidising agent : **Fluorine**
* The element of lowest ionisation energy : **Caesium**
* The element of highest ionisation energy : **Helium**
* The most electronegative element : **Fluorine**
* The element of highest electron gain enthalpy : **Chlorine**
* The group containing most electropositive metals : **Group 1**
* The group containing most electronegative metals : **Halogens Group 17**
* The group containing maximum number of gaseous elements : **Group 18**

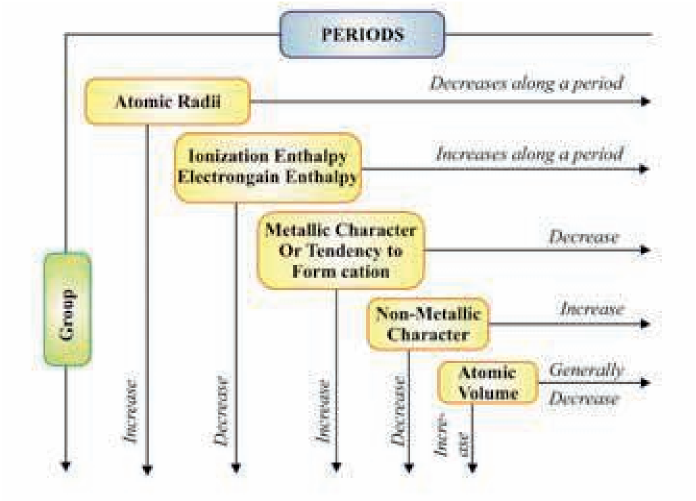
13. **Second period element**—Show different behaviour that I group **element**—Due to (a) small size (b) High electronegativity (C) High polarising power (d) absence of '*d*' orbital.

$Na_3[Al(OH)_6]$ exists but $Na[B(OH)_4]$ not exists.

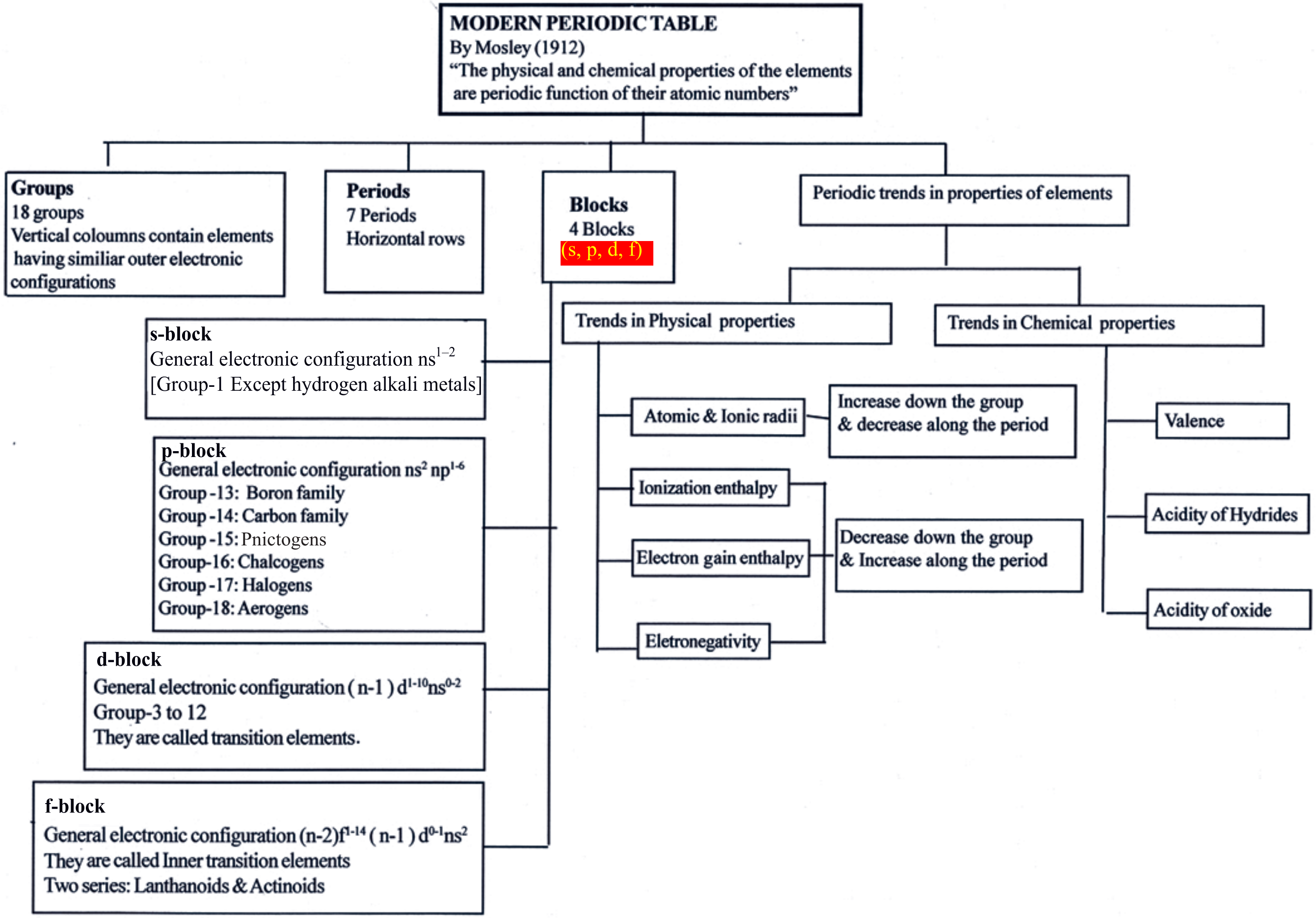
14. The similarities in properties of first member of a group to second member of just next higher group due to comparable atomic radius, nearly same polarising power of ions is known as **diagonal relationship.**



Elements with number of e^-		**in valance shell**
(a)	1, 2, 3	metals
(b)	4	metalloids
(c)	5, 6, 7	non-metals
(d)	8	noble gas



MIND MAP
CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES



MODERN PERIODIC TABLE
By Mosley (1912)
"The physical and chemical properties of the elements are periodic function of their atomic numbers"

Groups
18 groups
Vertical coloumns contain elements having similiar outer electronic configurations

Periods
7 Periods
Horizontal rows

Blocks
4 Blocks
(s, p, d, f)

s-block
General electronic configuration ns^{1-2}
[Group-1 Except hydrogen alkali metals]

p-block
General electronic configuration $ns^2\ np^{1-6}$
Group -13: Boron family
Group -14: Carbon family
Group -15: Pnictogens
Group-16: Chalcogens
Group -17: Halogens
Group-18: Aerogens

d-block
General electronic configuration $(n-1)\ d^{1-10}ns^{0-2}$
Group-3 to 12
They are called transition elements.

f-block
General electronic configuration $(n-2)f^{1-14}\ (n-1)\ d^{0-1}ns^2$
They are called Inner transition elements
Two series: Lanthanoids & Actinoids

Periodic trends in properties of elements

Trends in Physical properties
- Atomic & Ionic radii
- Ionization enthalpy
- Electron gain enthalpy
- Eletronegativity

Increase down the group & decrease along the period

Decrease down the group & Increase along the period

Trends in Chemical properties
- Valence
- Acidity of Hydrides
- Acidity of oxide

CASE BASED STUDY QUESTIONS

1. Read the passage given below and answer the following questions:

A period is a horizontal row in the periodic table. Although groups generally have more significant periodic trends, there are regions where horizontal trends are more significant than vertical group trends, such as the f-block, where the lanthanides and actinides form two substantial horizontal series of elements.

Elements in the same period show trends in atomic radius, ionization energy, electron affinity, and electronegativity. Moving left to right across a period, atomic radius usually decreases. This occurs because each successive element has an added proton and electron, which causes the electron to be drawn closer to the nucleus. This decrease in atomic radius also causes the ionization energy to increase when moving from left to right across a period. The more tightly bound an element is, the more energy is required to remove an electron. Electronegativity increases in the same manner as ionization energy because of the pull exerted on the electrons by the nucleus. Electron affinity also shows a slight trend across a period. Metals (left side of a period) generally have a lower electron affinity than non-metals (right side of a period), with the exception of the noble gases.

(Reference: https://en.wikipedia.org/wiki/Periodic_table)

The following questions are multiple choice questions. Choose the most appropriate answer:

(i) The atomic radii of Elements Z and X are compared. Element Z is having larger radius than Element X. (Both the elements do not have noble gas configuration and exist in same period in the periodic table)

Based on this you can say that the:

A) Element Z is located to the left side of Element X in the periodic table

B) Element Z is located to the right side of Element X in the periodic table

C) Element Z and X are probably in the same group

D) None of the above

(ii) In which of the following atoms is the 3s orbital closest to the nucleus?

A) Br B) Cl

C) I D) Same distance in all of these atoms

(iii) __________ have the lowest first ionization energies of the groups listed.

A) Alkali metals

B) Transition metals

C) Halogens

D) Noble gases

(iv) The correct order of electronegativity is

A) Cl > F > O > Br

B) F > O > Cl > Br

C) F > Cl > Br > O

D) O > F > Cl > Br

ANS:- I-A, II-D, III-A, IV-B

2. Read the passage given below and answer the following questions:

As the number of protons increase within a period (or row) of the periodic table, the first ionization energies of the transition-metal elements are relatively steady, while that for the main-group elements increases. The effective nuclear charge mirrors and may explain the periodic trends in the first ionization energies of the transition-metal and main-group elements. The differing periodic trends in the effective nuclear charge are due to a greater increase in shielding in the transition-metal elements than in the main-group elements. The difference in shielding is due to the entry of electrons into an inner-shell orbital for the transition-metal elements, while electrons enter an outer-shell orbital for the main-group elements.

(Reference: Paul S. Matsumoto J. Chem. Educ. 2005, 82, 11, 1660 Publication Date: November 1, 2005, Journal of the American Chemical Society)

Q.1. Why the first Ionisation energy of Be is greater than that of B?

Q.2. Why the ionisation of s-electron require more energy than ionisation of p-electron of the same shell.

Q.3. The first ionisation enthalpy of aluminium is lower than that of magnesium. Why?

Q.4. Why the first ionisation energies of transition metal elements are relatively steady?

MULTIPLE CHOICE QUESTIONS (MCQ)

1. According to modern periodic law, the physical and chemical properties of elements are the periodic functions of their ?

(a) Density
(b) Atomic Number
(c) Mass Number
(d) Atomic Mass

2. Highest electropositive element in the periodic table is

(a) Cs
(b) Rb
(c) K
(d) Na

3. The correct order of ionic radii of the species N^{3-}, O^{2-}, Na^{+} and F^{-} is

(a) $Na^{+} < F^{-} < O^{2-} < N^{3-}$
(b) $F^{-} < O^{2-} < N^{3} > Na^{+}$
(c) $O^{2-} < N^{3-} < F^{-} > Na^{+}$
(d) $N^{3-} < Na^{+} < F^{-} > O^{2-}$

4. The basic strength of the oxides follows the order

(a) $Al_2O_3 > MgO > Na_2O$
(b) $Al_2O_3 < MgO < Na_2O$
(c) $Na_2O_3 < MgO > Al_2O_3$
(d) $Al_2O_3 > MgO > Na_2O$

5. The correct order of the size of C, N, P, S follows the order

(a) $N < C < P < S$
(b) $C < N < S < P$
(c) $C < N < P < S$
(d) $N < C < S < P$

6. Which of the following oxide is most acidic?

(a) Na_2O
(b) Al_2O_3
(c) P_2O_5
(d) SO_3

7. Downward in a group, electropositive character of elements

(a) increases
(b) decreases
(c) remains same
(d) none of these

8. Element which has more negative electron gain enthalpy is

(a) F
(b) O
(c) Cl
(d) S

9. The electronegativity of the following elements increase in the order

(a) C, N, Si, P
(b) N, Si, C, P
(c) Si, P, C, N
(d) P, Si, N, C

10. The ionisation enthalpy of nitrogen is more than that of oxygen molecules because of
(a) greater attraction of electrons by the nucleus
(b) extra stability of the half filled p-orbitals
(c) smaller size of nitrogen
(d) more penetrating effect

Ans: 1. (b), 2. (a), 3. (a), 4. (b), 5. (d), 6. (d), 7. (a), 8. (c),
9. (c), 10. (b)

FILL IN THE BLANKS

1. Lightest metal in s-block elements is ______________ .
2. In the periodic table, horizontal rows are known as ______________ .
3. Elements of s-blocks and p-blocks are collectively called ____________.
4. Most electropositive elements belong to ____________ group.
5. Most electronegative elements belong to ____________ group.
6. The elements above atomic number 92 are called ____________.
7. The inner-transition elements belong to _________ block of the periodic table and are shown separately at the _________ of the periodic table.
8. An element having electronic configuration [Ar] $3d^5$, $4s^2$ belongs to _____________ block.
9. Ca^{2+} has smaller ionic radius than K^+ ion because it has ____________.
10. The maximum electronegativity is shown by _____________.
11 The maximum ionisation enthalpy is shown by ____________.
12 The cation is _________ and the anion is _________ than the parent atom.

Ans:
1. Lithium
2. periods
3. normal elements or representative elements
4. 1st
5. 17th
6. transuranic elements
7. F –, bottom
8. s – d
9. more protons
10. F –
11. He
12. smaller, bigger

TRUE AND FALSE TYPE QUESTIONS

Write true or false for the following statements

1. First ionisation enthalpy of Be is higher than B.
2. Every period of the periodic table (except first period) starts with a member of alkali metal.
3. The energy required during the removal of one electron from an atom is called its ionisation potential.
4. Flourine has more negative electron gain enthalpy than chlorine.
5 Mg^{2+} ion has smaller size than Mg.
6. Electronegativity of F is larger than that of Cl but electron gain enthalpy of Cl is larger than of F.
7. The decreasing order of electronegativity of F, O and N is F > O > N.
8. Group-18 contain maximum gaseous elements.
9. Al_2O_3 is an amphoteric oxide.
10. Helium has the highest ionisation enthalpy.

Ans: 1. (T) 2. (T) 3. (T) 4. (F) 5. (T)
6. (T) 7. (T) 8. (T) 9. (T) 10. (T)

MATCH THE COLUMNS

1.

	Column A		Column B		Column C
a.	Lightest element	i.	Caesium	p.	Is^{1}
b.	Lightest metal	ii.	Osmium	q.	[He] 2s^{1}
c.	Heaviest metal	iii.	Lithium	r.	[Xe] 6s^{1}
d.	Most reactive metal	iv.	Hydrogen	s.	d-block element

2.

	Column A		Column B
a.	Fluorine	i.	High negative electron gain enthalpy
b.	Helium	ii.	Most electropositive element
c.	Chlorine	iii.	Most electronegative element
d.	Caesium	iv.	Highest ionisation enthalpy

	Column C
p.	[Xe] 6s^{1}
q.	[He] 2s^{2} 2p^{5}
r.	Is^{2}
s.	[Ne] 3s^{2} 3p^{5}

3.

	Column A		**Column B**
a.	Na_2O	i.	Amphoteric oxide
b.	Cl_2O_7	ii.	Acidic oxide
c.	Al_2O_3	iii.	Neutral oxide
d.	CO	iv.	Basic oxide

4.

	Column A		**Column B**
a.	s & p-block	i.	Inner transition elements
b.	d-block	ii.	s-block elements
c.	f-block	iii.	Transition elements
d.	group-1 and group-2	iv.	Representative elements

Ans: 1. a. (iv). (p), b. (iii). (q), c. (ii). (s), d. (i). (r)
2. a. (iii). (q), b. (iv). (r), c. (i). (s), d. (ii). (p)
3. a.(iv), b.(ii), c.(i), d.(iii)
4. a.(iv), b.(iii), c.(i), d.(ii)

ASSERTION AND REASON TYPE QUESTIONS

Directions for Q. No.1-10

A Both Assertion & Reason are true and the reason is the correct explanation of the assertion.

B Both Assertion & Reason are true but the reason is not the correct explanation of the assertion.

C Assertion is true statement but Reason is false.

D Assertion is false but Reason is true.

1. Assertion : Ionic radius of Na^+ is smaller than Na
Reason : Effective nuclear charge of Na^+ is higher than Na

2. Assertion : First ionisation enthalpy of N is higher than O.
Reason : Extra stability of half filled up 2p subshell of N atom

3. Assertion : Electron gain enthalpy of Cl is more negative than F atom.
Reason : F is more electronegative than Cl atom.

4. Assertion : First ionisation enthalpy of Galium is higher than aluminium.
Reason : Weak sheliding effect of 3d subshell is Galium.

5. Assertion: Noble gases have positive electron gain enthalpy.
Reason: Noble gases have stable closed shell electronic configuration.

6. Assertion: F is more electronegative than Cl.
Reason: F has more electron affinity than Cl.

7. Assertion: The ionic size of O^{2-} is bigger than that of F^- ion.
Reason: O^{2-} and F^- are isoelectronic ions.

8. Assertion: The ionic radii follows the order: $I^- < I < I^+$.
Reason: Smaller the value of zeff, larger the size of the species.

9. Assertion: The first ionisation enthalpy of aluminium is lower than that of magnesium.
Reason: Ionic radius of aluminium is smaller than that of magnesium.

10. Assertion: First ionisation energy for nitrogen is higher than that of oxygen.
Reason: Across a period effective nuclear charge decreases.

Ans: 1. A 2. A 3. B 4. A 5.A 6.C 7.B 8.D 9.B 10.C

ONE WORD ANSWER TYPE QUESTIONS

1. Metals are placed on which side of modern periodic table?
2. Which block of modern periodic table represents inner transition elements?
3. Name a halogen which has most negative electron gain enthalpy value?
4. Which element is iso-electronic with Na^+ ? [**Ans.** Ne]
[Given atomic number of Sodium (Na) : 11]
5. An element is placed in 5th period and 3rd group what is its atomic number? [**Ans.** 39]
6. What is covalency of Al in $[AlCl_4]^-$? [**Ans.** 4]
7. Write the IUPAC Symbol for the element having atomic number 120. [**Ans.** Ubn]
8. Write the name of the group containing maximum number of gaseous elements.
9. Write the name of the subshell which show weakest sheilding effect.
10. Write the name of most electropositive element in the periodic table.
11. In what period and group will an element with Z = 118 will be present.

1-MARK QUESTIONS

1. Which pair of elements has similar properties?
   13, 31, 11 & 21
2. Name the element which exhibit diagonal relationship with Be.
   [**Ans.** 13, 31]
3. Which group elements are known as halogens?
4. The element with ns^2, np^5 configuration is non-metal or metal?
5. Define van der Waal's radius.
6. Write the outer shell configuration of atomic number 31. [**Ans.** $4s^2, 4p^1$]
7. Find the group number and period number of element having atomic number 52. [**Ans.** Period = 5th, Group = 16th]
8. Arrange O^{2-}, O^{-1}, O in decreasing radius (size). [**Ans.** $O^{2-} > O^{-1} > O$]
9. Why noble gas have bigger size than halogens?
10. Why first electron gain enthalpy of sulphur is more negative then oxygen?
11. Write general outer electronic configuration of 4f series elements.
    [**Ans.** $6s^2$, $5d^{0-1}$, $4f^{1 \text{ to } 14}$]
12. Write two isoelectronic species with Br (35). [**Ans.** Kr^+, Sc^-]
13. Show that 4th period can have maximum 18 elements in it.
14. Second I.E. is always more than first I.E., why?
15. Electronegativity of F > Cl > Br > I, why?
16. Arrange F and Cl in terms of increasing chemical reactivity?
17. Second I.E. of Na is more than second IE of Mg. Why?
18. I.E. for cation is more than neutral atom. Why?
19. Define diagonal relationship with the help of an example.
20. Out of O^- and O, which has more negative electron gain enthalpy?
21. Mention any two anomalous properties of second period elements.

2-MARKS QUESTIONS

1. Cations are smaller than their parent atom whereas anions are larger in size than their parent atom. Explain.

2. Ionisation energy of nitrogen is more than 'O' and 'C' both, why ?

3. First ionisation energy of boron is less than Be but size of Be is less than Boron. Why ?

4. Electron gain enthalpy of Mg is positive. Explain.

5. Define co-valency.

6. The reactivity of halogens decrease down the group but of alkali metals increases down the group. Why?

7. Name a halogen, a metal and a group13 element which are liquid at 30°C. [**Ans.** Br, Hg, Ga]

8. The reducing power of elements increases down the group but reverse is true for oxidising power along a period. Why ?

9. What is the formula of binary compound formed between :

(a) 1st element of I group and iodine ?

(b) 2nd element of II group and 1st element of 17th group ?

10. Arrange in the following in increasing order of property indicated:

(a) Size I, F, Cl, Br

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

11. Oxygen is more non-metallic than nitrogen but less than fluorine why ?

12. LiCl, LiBr, LiI are covalent as well as ionic why ?

13. $PbCl_2$ is more stable than $PbCl_4$. Why ? [**Ans.** Inert pair effect]

14. Magnesium and Lithium both form nitrides why?

15. Which has least I.E. [$3p^3$, $3p^6$, $2p^3$, $2p^6$]?

16. (a) I.E. of sulphur is lower than chlorine.

(b) Arrange the following in decreasing order of their electro-negativity:

F, O, N, Cl, C, H.

17. Element 'A' in group 17 (2nd period)

'B' in group 16 (2nd period)

'C' in group 15 (2nd period)

Arrange 'A', 'B' and 'C' in their decreasing order of electro-negativity and ionisation enthalpy.

18. Element 'A' 13 of group forms ionic compounds. Write the :

(a) Formula of its oxide.

(b) Arrange the following in their decreasing electro-positive character Mg, Na, Al, Si.

19. Write the atomic number of element placed diagonally to :

(a) Group 14, period 4 (b) Group 2, period 5

(c) Group 17, period 4

20. An element has outer shell electronic configuration $4s^2\ 4p^3$. Find :-

(a) The atomic number of element place next below it.

(b) Atomic number of next noble gas.

3-MARKS QUESTIONS

1. What is metallic radius, Covalent radius, van der waal's radius. Give one example for each.

2. Oxygen has first electron gain enthalpy exothermic while second endothermic still a large number of ionic oxides are formed. Why ?

3. In some properties Boron shows different properties with respect to rest of the member in the group. Justify.

4. Out of group 17, 18 and 1, predict:-

(a) Which has most negative first electron gain enthalpy ?

(b) Which shows most metallic behaviour ?

(c) Which has highly positive electron gain enthalpy?

5. What are (a) representative elements, (b) Transition elements, (c) Lanthanoid and actinoids. Give their positions in modern periodic table.

6. Why LiF, NaF, KF, RbF, CsF are ionic ? But LiF is less ionic than CsF.

7. (a) Why Ca has larger atomic radius than Al ?

(b) Why $2s^2$ electron is difficult to remove than $2p$ electron ?

8. (a) Why the compounds of group 17 with group 13 elements are more ionic and stable than with (group 1) elements? (b) Na_2O is more ionic than Li_2O. why?

9. Explain the following data :

Ionisation energy Cl < H < O < N < F.

10. IE_2 of 3rd period elements is as follows. Why ?

$Mg < Si < Al < P < S < Cl < Ar < Na$.

11. Account fot the following:

(a) Halogens have very high negative electron gain enthalpy

(b) The electron gain enthalpy of Cl (Z = 17)is more negative than that of Fluorine (Z = 9).

(c) Ionisation enthaply of Nitrogen (Z = 7) is more than oxygen (Z =8).

12. What are the d- block elements? Write any four properties of d - block elements and give their general outer electronic configuration.

13. Explain the following:

(a) Modern Periodic law

(b) Electro-negativity

(c) Shielding effect

14. Among the second period elements the actual ionisation enthalpies are in the order $Li < B < Be < C < O < N < F < Ne$. Explain why?

(i) Be has higher $(\Delta_i H)_1$ than B

(ii) O has lower $(\Delta_i H)_1$ than N and F ?

15. What do you understand by the isoelectronic species ? Name a species that will be isoelectronic with each of the following atoms or ions.

(i) F^- (ii) Ar (iii) Ca^{2+} (iv) Rb^+

16. (a) Show by a chemical reaction with water that Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide.

(b) Name a species that will be isoelectronic with each of the following atoms or ions, (i) F^- (ii) Ca^{2+}

17. The first ionisation enthalpy values (in $kJmol^{-1}$) of group-13 elements are:

B	Al	Ga	In	Tl
801	577	579	558	589

How would you explain this deviation from the general trend ?

18. The first (IE_1) and the second (IE_2) ionisation enthalpies (kJ mol^{-1}) of three elements are given below:

	I	II	III
IE_1	403	549	1142
IE_2	2640	1060	2080

Identify the element which is likely to be:-

(a) a non metal

(b) an alkali metal

(c) an alkaline earth metal

5-MARKS QUESTIONS

1. (A) Which of the following have same chemical properties :

(a) Atomic number 17, 53

(b) Atomic number 8, 52

(c) Both

(d) None

(B) Answer the following :

(i) B, Al, Ga (decreasing order of atomic radii).

(ii) C, S, N (decreasing order of (ΔegH_1)

(iii) Al forms amphoteric oxide. Why ?

(iv) Mg^{2+} ion is smaller than O^{2-} ion although both have the some electronic configuration.

2.

Element	$\Delta_i H^{\ominus}_1$	$\Delta_i H^{\ominus}_2$	ΔegH^0_1
I	1681	3374	– 328
II	1008	1846	– 295
III	2372	5251	+ 48

(a) The most reactive non-metal.

(b) The least reactive non-metal.

(c) The least reactive element. Give reasons also.

[**Ans.** (a) 1 (b) II (c) III]

UNIT TEST-I

Time allowed : 1 Hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. Which of the following show the weakest shielding effect ? (1)

   (a) s (b) p (c) d (d) f

2. Which has highest electronegativity ? (1)

   (a) Cl (b) O (c) N (d) S

3. Which pair of elements has similar properties? (1)

   13, 31, 11, 21

4. Write general outer electronic configuration of 4f series elements. (1)

5. Write the IUPAC symbol for the element having atomic number 120. (1)

6. (a) Explain why cation are smaller and anions larger in radii than their parent atoms? (2)

   (b) Explain why ionisation enthalpy of nitrogen is more than that of oxygen.

7. The first ionisation enthalpy values (in kJ mol^{-1}) of group-13 elements are : (2)

B	Al	Ga	In	Tl
801	577	579	558	589

   How would you explain this deviation from the general trend?

8. (a) Show by a chemical reaction with water than Na_2O is a basic oxide and Cl_2O_7 is an acidic oxide. (3)

   (b) Name a species that will be isoelectronic with each of the following atoms or ions. (i) F^- (ii) Ca^{2+}

9. Explain the following :

   (a) Shielding effect

   (b) Diagonal relationship

   (c) Anomalous behavior of second period elements.

10. (a) Alkali metals do not form dipositive ions. Why? (5)

    (b) What is the IUPAC name and symbol of the element having atomic number 117.

    (c) Are the oxidation state and covalency of Al in $[Al(H_2O)_6]^{2+}$ same?

    (d) Why are there fourteen elements in the Lanthanide series?

UNIT TEST-II

Time allowed : 1 Hour **Maximum Marks : 20**

General instructions :

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. In the P^{3-}, S^{2-} and Cl^{-} ions, the increasing order of size is (1)

   (a) Cl^{-}, S^{2-}, P^{3-} (b) P^{3-}, S^{2-}, Cl^{-}

   (c) S^{2-}, Cl^{-}, P^{3-} (d) S^{2-}, P^{3-}, Cl^{-}

2. The element with positive electron gain enthalpy is (1)

   (a) hydrogen (b) sodium (c) oxygen (d) neon

3. Write the IUPAC name and symbol for the element with atomic number 118. (1)

In following questions a statement of question followed by a statement of reason is given. Choose the correct answer out of the following choices :

(a) Both Assertion and Reason are true and Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are true but Reason is not the correct explanation of Assertion.

(c) Assertion is true but Reason is false.

(d) Both Assertion and Reason are false.

4. **Assertion :** Electron gain enthalpy becomes less negative as we go down a group. (1)

   Reason : Size of the atom increases on going down the group and the added electron would be farther from the nucleus.

5. **Assertion :** Boron has a smaller first ionisation enthalpy than Beryllium.(1)

   Reason : The penetration of 2s electron to the nucleus is more than 2p electron hence 2p electron is more shielded by the inner core than 2s electron.

6. Out of O and S, which has higher negative electron gain enthalpy and why? (2)

7. Assign the position of elements having outer electronic configuration :

   (i) $ns^2\ np^4$ for $n = 3$

   (ii) $(n-1)\ d^2\ ns^2$ for $n = 4$

8. Consider the element N, P, O and S and arrange them is order of : (3)
   (i) increasing 1st I.E.
   (ii) increasing negative electron gain enthalpy
   (iii) increasing non-metallic character

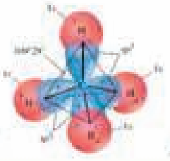
9. The first (IE_1) and second (IE_2) ionisation enthalpies ($kJmol^{-1}$) of three elements I, II and III are given below :

Element	IE_1	IE_2
I	403	2640
II	549	1060
III	1142	2080

   Identify the element which is likely to be
   (i) non-metal
   (ii) an alkali metal
   (ii) an alkaline earth metal (3)

10. (a) Lithium shows diagonal relationship with which element and why?
    (b) Among the elements of second period Li to Ne, pick out element:
       (i) with the highest 1st I.E.
       (ii) with the highest electronegativity
       (iii) with largest atomic radius
       (iv) most reactive non-metal (5)

*****



Chapter - 4

Chemical Bonding and Molecular Structure

FAST TRACK : QUICK REVISION

- **Kossel-Lewis Concept:** Atoms take part in chemical combination to complete octet in their valence shell. This is known as octet rule.
- **Limitation of Octet Rule:** The octet rule, though useful but has some exceptions e.g. BF_3, NO_2, PCl_5, SF_6 etc.
- **Lewis Symbol or Electron Dot Structure:** Representing valence electrons by dots placed around the letter symbol of the element.

Types of Chemical Bonds:

(i) **Covalent Bond:**
   (a) Formed by sharing of electrons.
   (b) It may be polar or nonpolar.
   (c) It is directional in nature.

(ii) **Ionic Bond:**
   (a) Formed by transfer of electrons.
   (b) Formation of ionic bond is favored by high lattice enthalpy, Low ionization enthalpy of metal atom and more negative electron gain enthalpy of nonmetal atom.
   (c) It is non directional in nature.

- **Formal Charge (F.C.):**
  (i) It is charge appeared on individual atom in covalent molecule.
  (ii) F.C. = (Total No. of valence electrons in free atom) – (Total No. of unshared electrons) – ½ (Total No. of shared electrons)
  Greater the F.C on atoms lesser the stability of that Lewis structure.

- **Lattice Enthalpy:** Energy released when one mole of a crystalline solid is formed from constituent gaseous ions.

Bond length:

(i) It is equilibrium distance between the nuclei of two bonded atoms in a molecule.

(ii) Greater the size of bonded atoms shorter the bond length.
e.g., $H-F < H-Cl < H-Br < H-I$

(iii) Greater the s character shorter the bond length.
e.g., $C_{sp^3}-H > C_{sp^2}-H > C_{sp}-H$

(iv) Bond length decreases with increase in bond order.
e.g., $C-C > C=C > C\equiv C$

- **Bond angle:**

(i) It is angle between the orbitals containing bonding electron pairs around central atom in a molecule or complex ion.

(ii) Greater the electronegativity of central atom larger the bond angle
e.g., $NH_3 > PH_3$

(iii) Greater the number of lone pair around central atom smaller the bond angle. **e.g.,** $CH_4 > NH_3 > H_2O$

- **Bond Enthalpy:**

(i) It is defined as amount of energy required to break one mole of bonds of a particular type between two atoms in gaseous state.

(ii) For diatomic molecules, Bond enthalpy $\propto$ Bond dissociation enthalpy

(iii) For polyatomic molecules, Bond enthalpy $\propto$ Average of all possible bond dissociation enthalpies.

(iv) Bond enthalpy α Bond order α 1/(Bond length)

- **Resonance:**

(i) According to the concept of resonance, whenever a single Lewis structure cannot describe a molecule accurately, a number of structures with similar energy, position of nuclei, bonding and non-bonding pairs of electrons are taken as canonical structures of the resonance hybrid which describes the molecule accurately.

(ii) Resonance averages the bond characteristics as a whole.

- **Partial ionic character of covalent bond A–B:**

$= 16(X_A - X_B) + 3.5(X_A - X_B)^2$,

where X_A and X_B are electro-negativities of A & B.

- **Partial covalent character in ionic bond (Fajan's rule):**
  - (i) Fajan's rule is used to predict partial covalent character in ionic bond.
  - (ii) Gretaer the polarizing power of cation and polarisability of anion greater the covalent character in ionic bond.
  - (iii) Polarising power of cation α Charge density [(Charge)/size)].
  - (iv) Polarisability of anion α size of anion.

- **Dipole moment:**
  - (i) Dipole moment (μ) = charge (Q) × distance of separation (d)
  - (ii) Unit: Debye (D), $1\text{D} = 3.33564 \times 10^{-30}$ Cm
  - (iii) Being vector quantity, dipole moment of polyatomic molecule is taken as the resultant of all the bond moments.
  - (iv) If $\mu = 0$, molecule is non polar or symmetric.
  - (v) If $\mu \neq 0$, molecule is polar or asymmetric.

- **Hydrogen bond:**
  - (i) It is dipole-dipole interaction between molecules in which 'H' atom is inserted between two highly electronegative elements i. e. F, O or N only.
  - (ii) Hydrogen bond may be intra-molecular (when present within single molecule) and intermolecular (when present b/w two same or different molecules).
  - (iii) Hydrogen bonds are stronger intermolecular forces than van der Waal forces.

- **Sigma (σ) and pi (π) bonds:**
  - (i) Sigma bond is formed by axial overlapping and pi bond is formed by sideways overlapping of atomic orbitals.
  - (ii) Sigma bond is stronger than pi bond due to greater extent of overlapping.
  - (iii) Single covalent bond = 1 σ bond

    Double covalent bond = 1 σ bond + 1 π bond

    Triple covalent bond = 1 σ bond + 2 π bond

- **VSEPR Theory:** (VSEPR = Valence Shell Electron Pair Repulsion): The shape of a molecule depends upon the number of valence shell electron pairs (lp and bp) around the central atom and magnitude of repulsive forces between them

  i.e., lp–lp > lp–bp > bp–bp

- **Hybridisation:**
  (i) It is the phenomena of mixing of atomic orbitals of nearly same energy to form the new orbitals of equal energy and identical shape.
  (ii) The new orbitals are called hybrid orbitals and determine the shape of molecules.
- **Molecular Orbital Theory (MOT):**
  (i) The intermixing of atomic orbitals of same symmetry to form bonding and antibonding molecular orbitals by addition and substraction of their wave functions is known as MO theory.
  (ii) The electrons are filled in molecular orbitals in order of their increasing energy.

  i.e., $\sigma 1s, \sigma^*1s, \sigma 2s, \sigma^*2s, \pi 2p_x = \pi 2p_y, \sigma 2p_z, \pi^*2p_x = \pi^*2p_y, \sigma^*2p_z$ (upto 14 electrons)

  $\sigma 1s, \sigma^*1s, \sigma 2s, \sigma^*2s, \sigma 2p_z, \pi 2p_x = \pi 2p_y, \pi^*2p_x = \pi^*2p_y, \sigma^*2p_z$ (For more than 14 electrons)

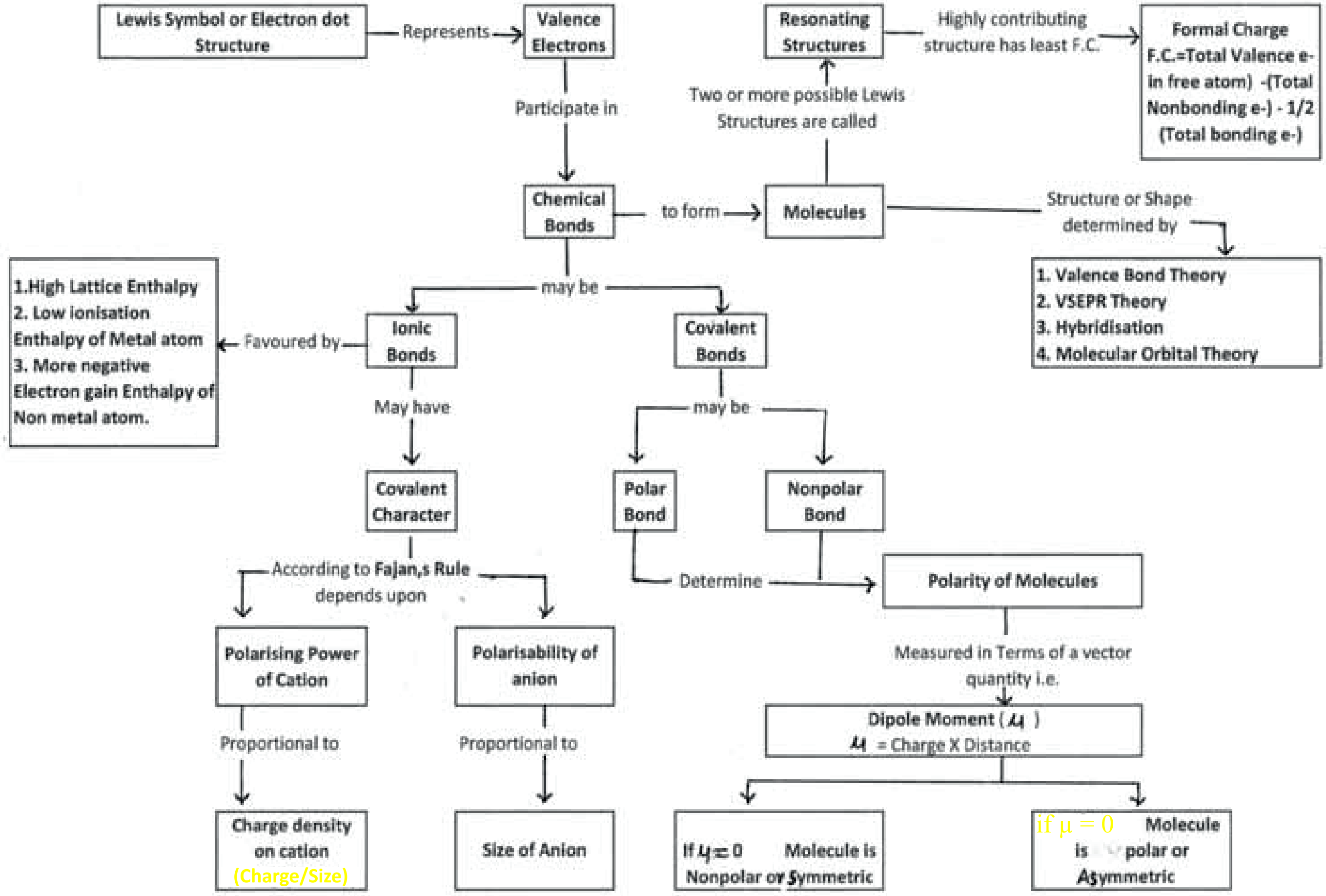
  (iii) Bond order = $1/2\,(N_b - N_a)$

  N_a = No of electrons in anti-bonding molecular orbitals

  N_b = No of electrons in bonding molecular orbitals

Total electron pairs	Bond pairs	Lone pairs	Type of hybridization	Geometry due to repulsion	Bond angle	Example
2	2	0	sp	Linear	180°	$BeCl_2$
3	3	0	sp^2	Non-polar Planar	120°	BF_3
3	2	1	sp^2	Angular	<120°	SO_2
4	4	0	sp^3	Tetrahedral	109°28′	CH_4
4	3	1	sp^3	Pyramidal	<109°28′	NH_3
4	2	2	sp^3	Bent	<109°28′	H_2O
5	5	0	sp^3d	Trigonal bipyramidal	120° & 90°	PCl_5
5	4	1	sp^3d	See Saw	<120 ° & <90°	SF_4
5	3	2	sp^3d	Bent T-shaped	<90°	ClF_3
5	2	3	sp^3d	Linear	180°	I_3^-
6	6	0	sp^3d^2	Octahedral	90°	SF_6
6	5	1	sp^3d^2	Square pyramidal	<90°	BrF_5
6	4	2	sp^3d^2	Square planar	90°	XeF_4
7	7	0	sp^3d^3	Pentagonal bipyramidal	90° & 72°	IF_7
7	6	1	sp^3d^3	Pentagonal pyramidal	<90° & <72°	
7	5	2	sp^3d^3	Pentagonal planar	72°	XeF_5^-

MIND MAP
CHEMICAL BONDING AND MOLECULAR STRUCTURE



CASE BASED STUDY-QUESTION

PASSAGE -1

Pauling introduced the concept of hybridisation. According to him the atomic orbitals combine to forms new set of equivalent orbitals known as hybrid orbitals. Unlike pure orbitals, the hybrid orbitals are used in bond formation .The phenomenon is known as hybridisation which can be defined as the process of intermixing of the orbitals of slightly different energies so as to redistribute their energies, resulting in the formation of new set of orbitals of equivalent energies and shape.

Source: NCERT

The following questions are multiple choice questions. Choose the most appropriate answer:

I. In SF_6 molecules , which sets of d-orbitals is involved ?

(A) dx^2-y^2, dz^2 (B) dz^2, dxy

(C) dxy, dyz (D) dx^2-y^2, dxy

II. In IF_7 molecule, which orbitals are involved

(A) dx^2-y^2, dz^2, dxy (B) dxy, dyz, dzx

(C) dx^2-y^2, dxy, dxz (D) dz^2, dyz, dzx

III. In PCl_5 molecule, d-orbitals involved is

(A) dxy (B) dyz

(C) dx^2-y^2 (D) dz^2

IV. Which of the following orbitals cannot undergoes hybridisation amongst themselves

(A) 3d, 4s (B) 2s, 2p

(C) 4s, 4d (D) 3s, 3p, 4s

ANS. : I-A, II-A, III-D, IV-D

PASSAGE -2

It can be said that covalent compound has partial ionic character due to the electronegativity difference of the two elements in a covalent bond .Dipole moment parameter is used to find the % of ionic character in the covalent compound. We can also say that ionic compounds do have some covalent character which can be explained qualitatively on the basis of the Fajan's Rule. According to Fajan rule a cation which has smaller size, high ionic charge must have large polarising power and the anion which has large size and high ionic charge must have large polarisibility. Both these factors help to develop covalent character in the molecule.

In these questions (Q. No V-VIII , a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(a) Assertion and reason both are correct statements and reason is correct explanation for assertion.

(b) Assertion and reason both are correct statements but reason is not correct explanation for assertion.

(c) Assertion is correct statement but reason is wrong statement.

(d) Assertion is wrong statement but reason is correct statement

V. ASSERTION:CO_2 molecule has zero dipole moment despite polar bond.

REASON: CO_2 molecule is linear in nature.

VI. ASSERTION: The dipole moment of Dichlorobenzene molecule can be zero.

REASON:The Dichlorobenzene must be 1, 4 Dichlorobenzene.

VII. ASSERTION: LiF has less covalent character than LiI

REASON: Halogen molecule are short of one electron to complete the octet.

VIII. ASSERTION: LiCl is more covalent than KCl

REASON: The polarising power of K is more than Li

ANS: V-A ,VI-A, VII-B, VIII-C

MULTIPLE CHOICE QUESTIONS (MCQ)

1. Which of the following molecules has both covalent and ionic bond

   (a) CH_3Cl (b) NH_4Cl (c) HCl (d) $BeCl_2$

2. What is the maximum number of water molecules that can attach with one water molecule through intermolecular hydrogen bonds?

   (a) 2 (b) 3 (c) 4 (d) 1

3. Which of the following molecules has maximum bond angle

   (a) NH_3 (b) CH_4 (c) H_2O (d) CO_2

4. Identify correct statement regarding NH_3 and BF_3

   (a) Both are Lewis acid
   (b) Both are iso structural
   (c) Both are Lewis base
   (d) Have different values of dipole moment

5. Identify the molecule having sideways overlapping of atomic orbitals

   (a) CH_4 (b) CO_2 (c) NH_3 (d) H_2O

6. Which of the following chemical species is most stable?

   (a) O_2 (b) O_2^+ (c) O_2^- (d) O_2^{2-}

7. Which of the following d orbitals involved in sp^3d hybridization?

   (a) d_{xy} (b) d_{xz} (c) $d_{x^2-y^2}$ (d) d_{z^2}

8. Which of the following molecule has net dipole moment?

   (a) CO_2 (b) H_2O (c) BF_3 (d) CH_4

9. Which of the following compound has highest covalent character

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

10. The shape of XeF_4 molecule according to VSEPR theory is

    (a) Square planar (b) Square pyramid
    (c) Tetrahedral (d) Pyramidal

Ans. 1.(b) 2.(c) 3.(d) 4.(d) 5.(b) 6.(b) 7.(d) 8.(b) 9.(d) 10.(a)

FILL IN THE BLANKS

(i) The energy required to completely separate one mole of solid ionic compound into gaseous constituent ions is called……………………..

(ii) Among alkali metal ions ………..ion has highest polarizing power.

(iii) According to molecular orbital theory molecules are said to be stable if the number of electrons in bonding molecular orbitals is …………… than the number of electrons in antibonding molecular orbitals.

(iv) Isoelectronic molecules and ions have identical…………………..

(v) In PCl_5 molecule the two equivalent axial P – Cl bonds are...............than three equivalent equatorial P – Cl bonds.

(vi) The state of hybridization of sulphur in SF_6 is……………………..

(vii) The maximum number of Hydrogen bonds formed by a single H_2O molecule is ....

(viii) A triple covalent bond consists of……sigma and…….pi bonds.

(ix) ………..bond is directional in nature.

(x) Atomic orbitals are………centric and molecular orbitals are………. centric.

Ans. (i) Lattice enthalpy (ii) Li^+ (iii) more (iv) bond order (v) longer (vi) sp^3d^2 (vii) 4 (viii) 1, 2 (ix) covalent (x) mono, poly

TRUE AND FALSE TYPE QUESTIONS

Write true or false for following statements:

(i) Energy of resonance hybrid is less as compared to the contributing canonical structures.

(ii) BeF_2 has more dipole moment than $BeCl_2$.

(iii) In water two O–H bond dissociation enthalpies are not identical.

(iv) Only the half filled orbitals of nearly same energy can participate in hybridization.

(v) No bond is purely ionic or purely covalent.

(vi) Chemical species having identical bond order have same bond dissociation enthalpies.

(vii) BF_3 is stronger Lewis acid than BCl_3.

(viii) Among alkali metal halides LiI has highest covalent character.

(ix) Resonating structures of a chemical species have no real existence.

(x) XeF_2 and ICl_2^- are iso structural.

Ans. (i) True (ii) False (iii) True (iv) False (v) True
(vi) False (vii) False (viii) True (ix) True (x) True

MATCH THE COLUMNS

I. Match the species in Column I with the geometry/shape in Column II and Hybridisation in Column III

S.N.	Column I	Column II	Column III
1.	BF_3	(a) Tetrahedral	(p) sp^3d^2
2.	ClF_3	(b) Trigonal Planer	(q) sp^3
3	NH_4^+	(c) Octahedral	(r) sp^3d
4.	SF_6	(d) Bent T-Shape	(s) sp^2

II. Match the species in Column I with the hybridsation in Column II and geometry/shape in Column III

S.N.	Column I	Column II	Column III
1.	BBr_3	(a) sp^3	(p) Square Planer
2.	H_2O	(b) sp^3d	(q) Trigonal Planer
3	PCl_5	(c) sp^2	(r) V-shape
4.	XeF_4	(d) sp^3d^2	(s) Trigonal Bipyramidal

ANS : MATCH-I

1. b, s 2. d, r 3. a, q 4. c, p

MATCH-II

1. c, q 2. a, r 3. b, s 4. d, p

ASSERTION AND REASON TYPE QUESTIONS

In the following questions a statement of assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below for each question:

(i) A and R both are correct, and R is correct explanation of A.

(ii) A and R both are correct, but R is not the correct explanation of A.

(iii) A is true but R is false.

(iv) A and R both are false.

1. Assertion (A): Among the two O–H bonds in H_2O molecule, the energy required to break the first O–H bond and the other O–H bond is the same.

   Reason (R): This is because the electronic environment around the oxygen is same after breakage of one O–H bond.

2. Assertion (A): Though the central atom of both NH_3 and H_2O molecules are sp^3 hybridised, yet H–N–H bond angle is greater than that of H–O–H.

   Reason (R): This is because nitrogen atom has one lone pair and oxygen atom has two lone pairs.

3. Assertion (A): SF_6 molecule is unstable.

   Reason (R): A stable molecule must have 8 electrons around the central atom. i.e. octet rule should be satisfied.

4. Assertion (A): Pi bond is never formed alone. It is formed along with a sigma bond

   Reason (R): Pi bond is formed by sideway overlap of p- orbitals only.

5. Assertion (A): Ionic compounds tend to be non-volatile.

   Reason (R): Ionic compounds are solid.

6. ASSERTION : Bonding molecular orbitals are more stable than Antibonding Molecular orbitals.

   REASON : Electrons placed in Bonding molecular orbitals tend to hold the nuclei more together as compared to electrons placed in Antibonding molecular orbitals.

7. ASSERTION : Ortho-Nitrophenol has higher boiling point than Para-Nitrophenol

   REASON : Intermolecular Hydrogen bonding occurs in Ortho-Nitrophenol as compared to intramolecular hydrogen bonding in Para-Nitro Phenol.

8. ASSERTION :When p_x orbital combine with p_y orbital then a sigma bond is produced.

   REASON : Atomic orbitals of similar symmetry can result in positive and negative overlap.

9. ASSERTION : BF_3 molecule has zero dipole moment.

   REASON : BF_3 molecule shape is trigonal planer and symmetrical.

10. ASSERTION : O-O bond length in O3 molecule is identical.

    REASON : Ozone molecule is angular in shape

Ans. 1. (iv) 2. (i) 3. (iv) 4. (iv) 5. (ii) 6.(i) 7. (iv) 8. (iv) 9. (ii) 10. (ii)

ONE WORD ANSWER TYPE QUESTIONS

1. Write the formal charge on central oxygen atom in O_3 molecule?
2. Write the shape of AB_2E_3 type molecule.
3. Name the property used to measure the degree of polarity.
4. Name the covalent bond formed by axial overlapping of atomic orbitals.
5. Out of p_x, p_y, p_z orbitals which p orbital takes part in sp hybridization?
6. Name the molecular orbital having energy greater than that of combining atomic orbitals.
7. Name the intermolecular forces responsible for liquid state of water.
8. Name the phenomenon used to describe a molecule whose single Lewis structure cannot describe it.
9. Name the geometry involved in sp^3d hybridization.
10. Name the molecular theory that can explain magnetic character of molecules.

Ans. 1. +1, 2. Linear, 3. Dipole moment, 4. Sigma bond, 5. p_z
6. Antibonding molecular orbital, 7. Hydrogen bond, 8. Resonance,
9. Trigonal bipyramid, 10. Molecular orbital theory

1-MARK QUESTIONS

1. Why noble gases exist in mono atomic form?
2. Write the Lewis structure of NO_2^-.
3. Why NH_3 and BF_3 have different shapes?
4. How many sigma and pi bonds are present in HCN molecule?
5. Why sigma bond is stronger than pi bond?
6. Explain why BeH_2 molecule has zero dipole moment although the Be–H bonds are polar?
7. Which has highest bond angle? NO_2, NO_2^-, NO_2^+
8. What is magnetic character of anion of KO_2?
9. Why do atoms combine?
10. What is the significance of Lewis Symbols?
11. Why density of water is maximum at 277K?
12. Give structure of BrF_5 according to VSEPR theory.
13. Why NH_3 is liquid and PH_3 is a gas?
14. Why KHF_2 exists but $KHCl_2$ and $KHBr_2$ does not?

    [**Ans.** HF….$F^{\ominus}$ hydrogen bonding].
15. Boiling point of p-nitrophenol is more than O-nitrophenol why?
16. How paramagnetic character of a compound is related to the no. of unpaired electrons?
17. Define the term bond length.
18. He_2 molecule does not exist. Give reason.
19. Why PCl_5 dissociates to give PCl_3 and Cl_2?
20. Write the state of hybridization of O in H_2O.
21. Predict the shape of ClF_3 according to VSEPR theory.
22. Why ice has less density than water?
23. Why the H–P–H bond angle in PH_3 is less than H–N–H bond angle in NH_3?
24. At room temperature H_2O exist as liquid while H_2S exist as gas. Give reason.
25. NH_3 has higher boiling point than PH_3. Give reason.
26. Identify the chemical species having identical bond order: O_2^{2+}, N_2, O_2, O_2^{2-}.

2-MARKS QUESTIONS

1. What is Octet rule? What are its limitations?
2. The enthalpy needed to break the two O–H bonds in water are as follows:

   $H_2O\,(g) \rightarrow H\,(g) + O{-}H(g) \qquad \Delta_a H_1^0 = 493\ \text{kJ mol}^{-1}$

   $O{-}H\,(g) \rightarrow H\,(g) + O(g) \qquad \Delta_a H_1^0 = 424\ \text{kJ mol}^{-1}$

   What is the average bond enthalpy of H_2O?
3. Write two points of difference between sigma and pi bond.
4. Define Hydrogen bond. Is it weaker or stronger than van der Waal forces?
5. Define dipole moment. Give its significance.
6. Give applications of dipole moment.
7. Which is more polar and why, CO_2 or N_2O?
8. Discuss the partial ionic character of covalent bond by taking an example.
9. Draw the resonating structures of O_3 and calculate formal charges on each O atom.
10. O-Nitrophenol is steam volatile while p-Nitrophenol is not. Give reason.
11. Define bond enthalpy. Why the bond enthalpy of F_2 is less than that of Cl_2?
12. Define resonance. Draw resonating structures of CO_2.
13. Assign reason for the following;
    (i) NH_3 is freely soluble in water while PH_3 is not.
    (ii) B_2 is paramagnetic while C_2 is not.
14. Out of NH_3 and NF_3 which is more polar. Explain with the help of dipole moment.
15. N_2 is diamagnetic while O_2 is paramagnetic. Explain on the basis of Molecular orbital theory.
16. H_2^+ and H_2^- have same bond order. Which is more stable?
17. Differentiate between bonding and anti bonding molecular orbitals.
18. Discuss the conditions for the combination of atomic orbitals to form molecular orbitals.

19. Although Chlorine (EN = 3.2) is more electronegative than Nitrogen (EN = 3.0), yet chlorine does not form hydrogen bond while nitrogen does. Give reason. (Ans: larger atomic size of Cl).

20. ClF_3 is T shaped but BF_3 is planar. Explain.

21. $N(SiH_3)_3$ and $N(CH_3)_3$ are not isostructural. Give reason.

22. Draw molecular orbital diagram for N_2^+ molecule.

23. HCl is a covalent compound but it ionises in the solution?

24. The molecule of CO_2 is linear whereas that of $SnCl_2$ is angular why?

25. Arrange the following in the order of property indicated for each set:
   (i) O_2, O_2^+, O_2^-, O_2^{2-} (increasing stability)
   (ii) LiCl, NaCl, KCl, RbCl (increasing covalent character)
   (iii) NO_2, NO_2^+, NO_2^- (decreasing bond angle)
   (iv) H–F, H–Cl, H–Br, H–I (increasing bond dissociation enthalpy)

26. Arrange the following in the order of property indicated for each set:
   (i) H_2O, NH_3, H_2S, HF (increasing polar character)
   (ii) HF, HCl, HBr, HI (decreasing dipole moment)
   (iii) NO_3^-, NO_2^-, NO (decreasing 's' character of hybridization)
   (iv) $BeCl_2$, BCl_3, CCl_4, PCl_3 (increasing bond angle)

3-MARKS QUESTIONS

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

2. 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

3. What is meant by hybridization? Describe the shape of sp, sp^2 and sp^3 hybridised orbitals.

4. Define bond order. Calculate the bond order in N_2 and O_2 molecules.
5. Give molecular orbital energy level diagram of O_2^{2-}. Write its electronic configuration,magnetic behaviour and bond order.
6. Which of the following in each pair has larger bond angle
   (i) CO_2 , BF_3 (ii) H_2O, H_2S (iii) CH_4, C_2H_2
7. What is meant by resonance? Draw the resonating structures of carbonate ion and explain why all the C–O bond lengths are identical in carbonate ion?
8. Compare relative stability of following species and predict their magnetic properties:
   O_2, O_2^+, O_2^- (superoxide), O_2^{2-} (peroxide)
9. Draw the Lewis structure of the species as mentioned BF_3, SF_6, NO_2:
   (i) In which the central atom has incomplete octet.
   (ii) In which the central atom has an expanded octet,
   (iii) An odd electron molecule is formed.
10. Explain the structure of PCl_5 according to hybridization. Why all P–Cl bonds lengths are not equivalent in PCl_5?

5-MARKS QUESTIONS

1. Give reasons for the following:
   (a) NH_3 has higher boiling point than PH_3.
   (b) Ionic compounds do not conduct electricity in solid state.
   (c) LiCl is more covalent than KCl.
   (d) NH_3 is more polar than NF_3.
   (e) H_2O has bent structure.
2. (a) Define the term bond dissociation enthalpy. How is it related to bond order?
   (b) Explain why N_2 has greater bond dissociation enthalpy than N_2^+ while O_2 has lesser bond dissociation enthalpy than O_2^+?
3. Draw the shape of following molecules according to VSEPR theory;
   XeO_3, XeF_2, $XeOF_4$, SF_4, XeF_4

HOTS QUESTIONS

1. The bond angle of H_2O is 104.5° while that of F_2O is 102°. Explain why?

   Solution: The bond pair of electrons are drawn more towards F in F_2O, whereas in H_2O it is drawn towards O. So bp – bp repulsion in H_2O is greater than that in F_2O.

2. Anhydrous $AlCl_3$ is covalent. From the data given below, predict whether it would remain covalent or become ionic in aqueous solution.

   $\Delta_i H\ (AlCl_3) = 5137\ \text{kJ mol}^{-1}$, $\Delta_{hyd}\ H(Al^{3+}) = -4665\ \text{kJ mol}^{-1}$,

   $\Delta_{hyd}\ H(Cl^-) = -381\ \text{kJ mol}^{-1}$.

   Solution: Total energy released = $1\Delta_{hyd}\ H(Al^{3+}) + 3\Delta_{hyd}\ H(Cl^-)$

   $$= [(-4665) + (3 \times -381)]\ \text{kJ mol}^{-1} = -5808\ \text{kJ mol}^{-1}$$

   Total energy required = $\Delta_i\ H(AlCl_3) = 5137\ \text{kJ mol}^{-1}$

   Since energy released is greater than the energy required, the compound will ionize in aqueous solution.

3. The dipole moment of HCl is 1.03 D, and the bond length is 127 pm. Calculate the percent ionic character of HCl molecule.

   Solution: $\mu_{cal} = Q \times r = (1.6\times10^{-19}C) \times (127\times10^{-12}m) = 2.032\times10^{-29}C\ m$

   $$= (2.032 \times 10^{-29}C\ m) \times \frac{1D}{3.336\times10^{-30}Cm} = 6.09\ D$$

   $$\%\ \text{ionic character} = \frac{\mu_{obs.}}{\mu_{cal}} \times 100 = \frac{1.03D}{6.09D}\times 100 = 16.9\%$$

UNIT TEST-I

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.
(ii) Maximum marks carried by each question are indicated against it.

---

1. Identify the molecule having sideways overlapping of atomic orbitals [1]
   (a) CH_4 (b) CO_2 (c) NH_3 (d) H_2O

2. The shape of XeF_4 molecule according to VSEPR theory is [1]
   (a) Square planar (b) Square pyramid
   (c) Tetrahedral (d) Pyramidal

3. Write the Lewis structure of NO_2^-. [1]

4. Which has highest bond angle? NO_2, NO_2^-, NO_2^+ [1]

5. Draw the resonating structures of CO_2. [1]

6. The enthalpy needed to break the two O–H bonds in water are as follows:

$$H_2O(g) \longrightarrow H(g) + O{-}H(g) \qquad \Delta_a H_1^0 = 493\ \text{kJ mol}^{-1}$$

$$O{-}H\ (g) \longrightarrow H(g) + O(g) \qquad \Delta_a H_2^0 = 424\ \text{kJ mol}^{-1}$$

   What is the average bond enthalpy of H_2O? [2]

7. Out of NH_3 and NF_3 which is more polar. Explain with the help of dipole moment. [2]

8. Compare relative stability of following species and predict their magnetic properties: O_2, O_2^+, O_2^- (superoxide), O_2^{2-} (peroxide) [3]

9. Explain the structure of PCl_5 according to hybridization. Why all P–Cl bonds lengths are not equivalent in PCl_5? [3]

10. (i) N_2 is diamagnetic while O_2 is paramagnetic. Explain on the basis of Molecular orbital theory. [2]

    (ii) Give reasons for the following: [3]
    (a) NH_3 has higher boiling point than PH_3.
    (b) Ionic compounds do not conduct electricity in solid state.
    (c) LiCl is more covalent than KCl.

UNIT TEST-II

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.
(ii) Maximum marks carried by each question are indicated against it.

---

1. Identify the molecule in which carbon has 'sp' hybridisation. [1]

   (a) CO_2 (b) CH_4 (c) C_2H_4 (d) C_2H_2

2. The shape of the molecule SF_4 is [1]

   (a) Bent (b) See-saw
   (c) Tetrahedral (d) Square Planer

3. Write the Lewis structure of CO_3^{2-} [1]

In following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices:

(a) Assertion and Reason both are correct statements and Reason is correct explanation for assertion.

(b) Assertion and Reason both are correct statement but Reason is not correct explanation for assertion.

(c) Assertion is correct statement but Reason is wrong statement.

(d) Assertion is wrong statement but Reason is correct statement.

4. **Assertion :** Pie (π) bond are directional in nature. [1]

   Reason : Sigma bond are formed by axial approach of atomic orbitals.

5. **Assertion :** Boiling point of water is higher than H_2S. [1]

   Reason : Hydrogen bonding is feasible in Water but in H_2S there is no Hydrogen bonding.

6. Why dipole moment of BF_3 is zero but for PCl_3 it is non zero? [2]

7. Which one LiF or LiI is more ionic and why? [2]

8. Explain with the help of labeled diagram the Valence Bond Theory for formation of H_2 molecule. [3]

9. Explain the Octet rule with relevant example. Write two limitation of Octet rule also. [3]

10. Define Hybridisation. Write the salient features of hybridisation. Explain the hybridisation in SF_6 molecule with relevant diagram. [5]

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Chapter - 5

Chemical Thermodynamics

FAST TRACK : QUICK REVISION

- **System:** Specific part of universe in which observations are made.
- **Surroundings**: Everything which surrounds the system.
- **Types of the System**:

  Open System: Exchange both matter and energy with the surroundings. For example – Reactant in an open test tube.

  Closed System: Exchange energy but not matter with the surroundings. For example – Reactants in a closed vessel.

  Isolated System: Neither exchange energy nor matter with the surroundings. For example – Reactants in a thermos flask.

- **Thermodynamic Processes:**
  - (i) Isothermal Process: $\Delta T = 0$
  - (ii) Adiabatic process: $\Delta q = 0$
  - (iii) Isobaric process: $\Delta P = 0$
  - (iv) Isochoric process: $\Delta V = 0$
  - (v) Cyclic process: $\Delta U = 0$
  - (vi) Reversible process: Process which proceeds infinitely slowly by a series of equilibrium steps.
  - (vii) Irreversible process: Process which proceeds rapidly and the system does not have chance to achieve equilibrium.
- **Extensive Properties:** Properties which depend upon the quantity or size of matter present in the system. For example – mass, volume, internal energy, enthalpy, heat capacity, work etc.

- **Intensive Properties**: Properties which do not depend upon the quantity or size of matter present in the system. For example – temperature, density, pressure, surface tension, viscosity, refractive index, boiling point, melting point etc.
- **State Functions**: The variables of functions whose value depend only on the state of a system or they are path independent.

  For example – pressure (P), volume (V), temperature (T), enthalpy (H), free energy (G), internal energy (U), entropy (S), amount (n) etc.
- **Internal Energy**: It is the sum of all kind of energies possessed by the system.
- First Law of Thermodynamics: "The energy of an isolated system is constant."

  Mathematical Form: $\Delta U = q + w$
- **Sign Conventions for Heat (q) and Work (w):**

  (i) $W = +$ ve , if work is done on system

  (ii) $W = -$ ve , if work is done by system

  (iii) $q = +$ ve , if heat is absorbed by the system

  (iv) $q = -$ ve , if heat is evolved by the system
- **Work of Expansion/ compression:** $\omega = -P_{ext}(V_f - V_i)$
- **Work done in Isothermal Reversible Expansion of an Ideal Gas**:

  $$\omega_{rev} = -2.303\ nRT \log V_f / V_i$$

  or $$\omega_{rev} = -2.303\ nRT \log P_i / P_f$$
- **Significance of ΔH & ΔU**: $\Delta H = q_p$ and $\Delta U = q_v$
- **Relation between ΔH & ΔU**: $\Delta H = \Delta U + (n_p - n_r)RT$ for gaseous reaction

  (i) $\Delta H = \Delta U$ if $(n_p - n_r)$ is zero; *e.g.* $H_2\,(g) + I_2\,(g) \longrightarrow 2\,HI(g)$

  (ii) $\Delta H > \Delta U$ if $(n_p - n_r)$ is positive; *e.g.* $PCl_5\,(g) \longrightarrow PCl_3\,(g) + Cl_2\,(g)$

  (iii) $\Delta H < \Delta U$ if $(n_p - n_r)$ is negative; *e.g.* $N_2\,(g) + 3H_2\,(g) \longrightarrow 2NH_3\,(g)$
- **Heat Capacity (C):** Amount of heat required to raise the temperature of a substance by 1°C or 1 K.

  $$q = C\,\Delta T$$

- **Specific Heat Capacity (C_s):** Amount of heat required to raise the temperature of 1 g of a substance by 1°C or 1 K.

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

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

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

- **Standard State of a Substance:** The standard state of a substance at a specified temperature is its pure form at 1 bar.

- **Standard Enthalpy of Formation($\Delta_f H^o$):** Enthalpy change accompanying the formation of one mole of a substance from its constituent elements under standard condition of temperature (normally 298 K) and pressure (1bar).
  - $\Delta_f H^0$ of an element in standard state is taken as zero.
  - Compounds with –ve value of $\Delta_f H^0$ are more stable than their constituents.
  - $\Delta_r H^0 = \sum a_i \Delta_f H^0$ (products) $- \sum b_i \Delta_f H^0$ (reactants) ; Where 'a' and 'b' are coefficients of products and reactants in balanced equation.

- **Standard Enthalpy of Combustion($\Delta_c H^0$):** Enthalpy change accompanying the complete combustion of one mole of a substance under standard conditions (298 K, 1bar)

- **Hess's Law of Constant Heat Summation:** The total enthalpy change of a reaction remains same whether it takes place in one step or in several steps.

- **Bond Dissociation Enthalpy:** Enthalpy change when one mole of a gaseous covalent bond is broken to form products in gas phase.

  For example, $Cl_2(g) \longrightarrow 2Cl(g); \quad \Delta_{Cl-Cl} H^0 = 242 \text{ kJ mol}^{-1}$
  - For diatomic gaseous molecules; Bond enthalpy = Bond dissociation Enthalpy = Atomization Enthalpy

- For Polyatomic gaseous molecules; Bond Enthalpy = Average of the bond dissociation enthalpies of the bonds of the same type.
- $\Delta_r H^0 = \sum\Delta_{bond} H^0$ (Reactants) – $\sum\Delta_{bond} H^0$ (Products)

◆ **Spontaneous Reaction:** A reaction which can take place either an its own or under some initiation.

◆ **Entropy(S):** It is measure of degree of randomness or disorder of a system.

- $\Delta S_{sys} = \frac{(q_{rev})_{sys}}{\Delta T} = \frac{\Delta H_{sys}}{\Delta T}$
- Unit of Entropy = $JK^{-1}\ mol^{-1}$

◆ **Second Law of Thermodynamics:** For all the spontaneous processes totally entropy change must be positive.

$$\Delta S_{total} = \Delta S_{sys} + \Delta S_{surr} > 0$$

◆ **Gibbs Helmholtz Equation for determination of Spontaneity:**

$$\Delta G = \Delta H - T\Delta S$$

(i) If ΔG = –ve, the process is spontaneous

(ii) If ΔG = +ve, the process is non-spontaneous

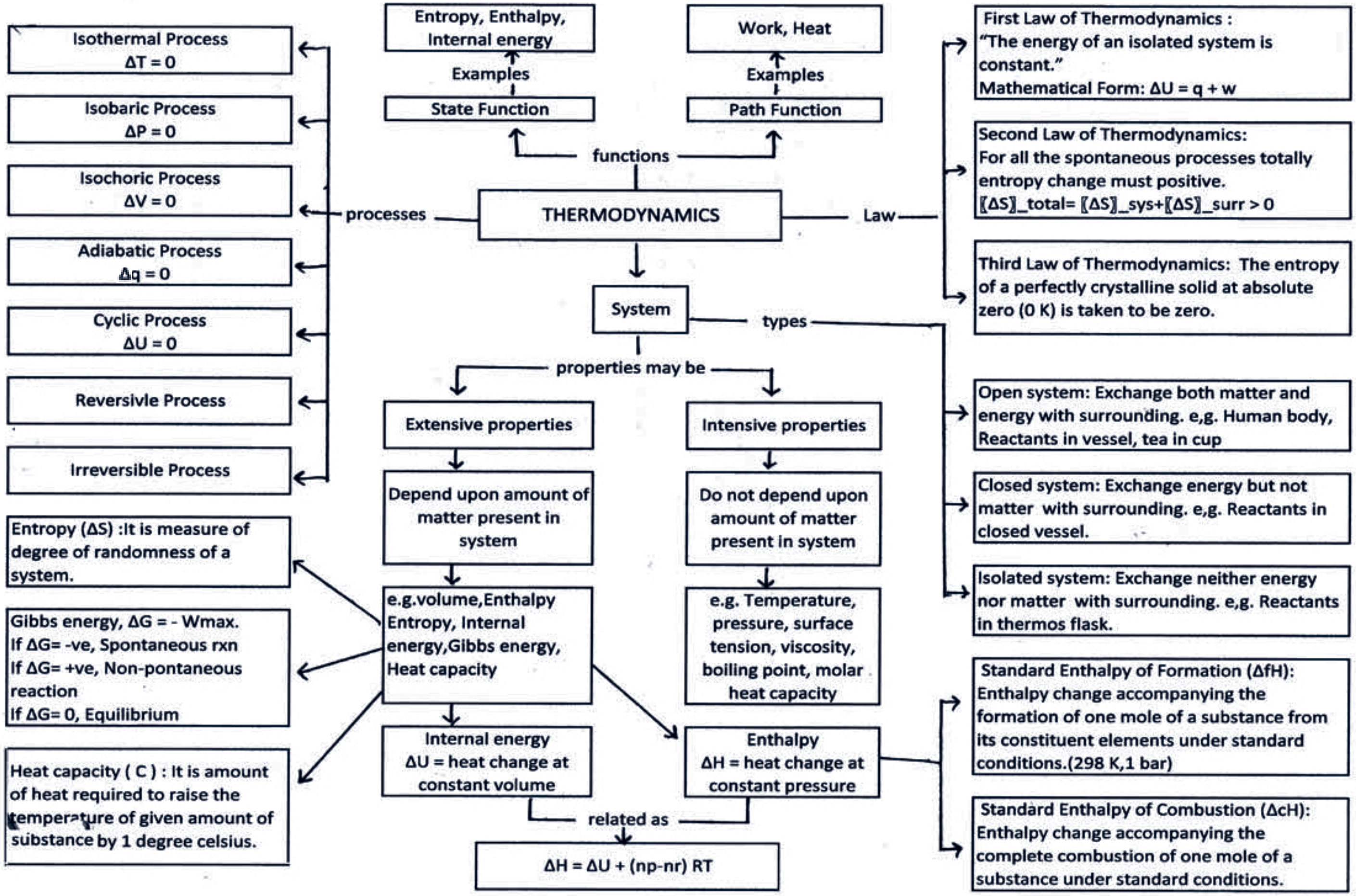
(iii) If $\Delta G = 0$, the process is in equilibrium

◆ **Relation between Gibbs Energy Change and Equilibrium Constant:**

$$\Delta G^o = -2.303\ RT\ \log K_c$$

◆ **Third Law of Thermodynamics:** The entropy of a perfectly crystalline solid at absolute zero (0 K) is taken to be zero.

MIND MAP : CHEMICAL THERMODYNAMICS



THERMODYNAMICS

processes

Isothermal Process $\Delta T = 0$

Isobaric Process $\Delta P = 0$

Isochoric Process $\Delta V = 0$

Adiabatic Process $\Delta q = 0$

Cyclic Process $\Delta U = 0$

Reversivle Process

Irreversible Process

functions

State Function

Examples

Entropy, Enthalpy, Internal energy

Path Function

Examples

Work, Heat

Law

First Law of Thermodynamics : "The energy of an isolated system is constant." Mathematical Form: $\Delta U = q + w$

Second Law of Thermodynamics: For all the spontaneous processes totally entropy change must positive. $[\![\Delta S]\!]_{total} = [\![\Delta S]\!]_{sys} + [\![\Delta S]\!]_{surr} > 0$

Third Law of Thermodynamics: The entropy of a perfectly crystalline solid at absolute zero (0 K) is taken to be zero.

System

types

Open system: Exchange both matter and energy with surrounding. e,g. Human body, Reactants in vessel, tea in cup

Closed system: Exchange energy but not matter with surrounding. e,g. Reactants in closed vessel.

Isolated system: Exchange neither energy nor matter with surrounding. e,g. Reactants in thermos flask.

properties may be

Extensive properties

Depend upon amount of matter present in system

e.g.volume,Enthalpy Entropy, Internal energy,Gibbs energy, Heat capacity

Entropy (ΔS) :It is measure of degree of randomness of a system.

Gibbs energy, $\Delta G = - W_{max}$.
If ΔG= -ve, Spontaneous rxn
If ΔG= +ve, Non-pontaneous reaction
If ΔG= 0, Equilibrium

Heat capacity (C) : It is amount of heat required to raise the temperature of given amount of substance by 1 degree celsius.

Internal energy ΔU = heat change at constant volume

Intensive properties

Do not depend upon amount of matter present in system

e.g. Temperature, pressure, surface tension, viscosity, boiling point, molar heat capacity

Enthalpy ΔH = heat change at constant pressure

Standard Enthalpy of Formation (ΔfH): Enthalpy change accompanying the formation of one mole of a substance from its constituent elements under standard conditions.(298 K,1 bar)

Standard Enthalpy of Combustion (ΔcH): Enthalpy change accompanying the complete combustion of one mole of a substance under standard conditions.

related as

$\Delta H = \Delta U + (np-nr) RT$

CASE BASED STUDIES - QUESTIONS

PASSAGE -1

Heat of neutralization is defined as amount of heat released when one gram equivalent of a strong base reacts with one gram equivalent of strong acid. The heat of neutralization comes out to be as follows

H^+ (aq.) + OH^- (aq.) → H_2O(aq.) , ΔH(Neu.) = 'X'

Since the basic reaction for neutralization reaction is same for all acids and bases, hence enthalpy of neutralization for strong acid and base is always same i.e., –57.1 kj. However, if any of the acid or base or both are weak then heat relased will be less. It is because some of heat released is utilized to ionize weak acid or base or both. In case of HF (weak acid), enthalpy of neutralization is more than that of strong acid and strong base due to high enthalpy of hydration of small size of ion.

Answer the following question

1. Enthalpy of neutralization of HCN against NaOH is – 51.4 kj. What is enthalpy of ionization of HCN?
2. If heat of neutralization of HA, HB, HC, HD acids follows the order HA > HB > HC > HD then arrange them in increasing order of acidic strength.
3. What will be heat change for complete neutralization of strong acid $R(OH)_2$ by HCI.
4. What will heat released when equal volume of an aqueous solution of 0.5M HNO_3 is mixed with 0.3 M NaOH(aq)

   OR

   What will be heat released when 100ml of 0.2 M HCl is mixed 200ml of 0.2M KOH.

Ans: 1. 5.7 KJ 2. HA > HB > HC > HD 3. –114.2 KJ or 1.142 KJ

PASSAGE -2

According to first law of thermodynamics, the mathematical expression is as follows :

$$\Delta U = Q + W$$

Where ΔU is the change in internal energy observed when a 'Q' amount of heat is involved and 'W' amount of work is associated when the system moves from state 'A' to state 'B'. The sign of 'Q' and 'W' is taken such that ΔU is measures accurately .

Answer the following question:

1. What will be work done for isochoric process?
2. What is relationship between internal energy change and work done for an adiabatic process?
3. Define internal energy change?
4. A sample of gas present in a cylinder fitted with a frictionless pistion, expands against a constant pressure of 1 atm from volume of 5L to 15L. During the process it absorb 500J energy from surrounding. Calculate change in internal energy.

   Or

   60 KJ heat is supplied to a system at constant volume and temperature rises from 20°C to 24°C. Calculate the change in internal energy, work done and q.

ANS: 1. 0 2. $\Delta E = w$ 3. Definition 4. –513 J or $\Delta E = 60$ J, $w = 0$, $q = 60$J

MULTIPLE CHOICE QUESTIONS (MCQ)

1. Which one of the following thermodynamic quantities is not a state function?

   (a) Gibbs free energy (b) Enthalpy

   (c) Entropy (d) Work

2. All of the following have a standard heat of formation value of zero at 25°C and 1.0 atm except:

   (a) N_2 (g) (b) Fe(s)

   (c) Ne (g) (d) H(g)

3. For the following reaction at 25°C, $\Delta H° = +115$ kJ and $\Delta S° = +125$ J/K. Calculate $\Delta G°$ for the reaction at 25°C :

   $$SBr_4(g) \longrightarrow S(g) + 2Br_2(l)$$

   (a) +152 kJ (b) –56.7 kJ

   (c) +77.8 kJ. (d) +37.1 kJ

4. Calculate $\Delta_r H^0$ for the following reaction at 25°C :

	$Fe_3O_4(s)$	$+ CO(g)$	$\longrightarrow 3FeO(s)$	$+ CO_2(g)$
ΔH_f^o (kJ/mol)	–1118	–110.5	–272	–393.5

   (a) –263 kJ (b) 54 kJ

   (c) 19 kJ. (d) –50 kJ

5. A system suffers an increase in internal energy of 80 J and at the same time has 50 J of work done on it. What is the heat change of the system?

   (a) +130 J (b) +30 J

   (c) –130 J (d) –30 J

6. The ΔH^0 for the following reaction at 298 K is –36.4 kJ.

   $$1/2\ H_2\ (g) + 1/2\ Br_2\ (l) \longrightarrow HBr(g)$$

   Calculate ΔU^0 at 298 K. The universal gas constant, R, is 8.314 J/mol K.

   (a) –35.2 kJ (b) +35.2 kJ

   (c) –36.4 kJ (d) –37.6 kJ.

7. For which of the following reactions would the $\Delta H°$ for the reaction be labeled ΔH_f^0 ?

   (a) $Al(s) + 3/2\ H_2\ (g) + 3/2\ O_2\ (g) \longrightarrow Al(OH)_3(s)$

   (b) $PCl_3\ (g) + 1/2\ O_2\ (g) \longrightarrow POCl_3\ (g)$

(c) $1/2\ N_2O\ (g) + 1/4\ O_2\ (g) \longrightarrow NO\ (g)$

(d) $CaO(s) + SO_2\ (g) \longrightarrow CaSO_3(s)$

8. Which statement is true for reaction? $2H_2\ (g) + O_2\ (g) \longrightarrow 2H_2O\ (g)$

(a) $\Delta S = +ve$ (b) $\Delta H > \Delta U$

(c) $\Delta H < \Delta U$ (d) $\Delta H = \Delta U$

9. The heat of combustion of yellow phosphorous is –9.91 KJ and the red phosphorous is –8.78 KJ. The heat of transition of yellow phosphorous to red phosphorous is :

(a) –9.91 kJ (b) –8.78 kJ

(c) –9.34 kJ (d) –1.13 kJ

10. Entropy of universe is :

(a) Increasing (b) decreasing

(c) Constant (d) None of these

11. Which is state function?

(a) q (b) w

(c) q + w (d) None of these

12. According to second law of thermodynamics

(a) $\Delta S_{total} = +ve$ (b) $\Delta S_{total} = -ve$

(c) $\Delta S_{system} = +ve$ (d) $\Delta S_{system} = -ve$

Ans: 1.(d), 2.(d), 3.(c), 4.(c), 5.(b), 6.(d), 7.(a),
8. (c), 9.(a), 10.(c), 11.(c), 12.(a)

FILL IN THE BLANKS

(i) ……… is a measure of the degree of randomness or disorder of a system.

(ii) A process which can take place either on its own or under some initiation is known as ………………..

(iii) For evaporation of water the sign of ΔH is…….. and sign of ΔS is……..

(iv) The entropy of a perfectly crystalline solid is zero at ………..

(v) The heat energy exchanged between the system and surroundings at constant temperature and pressure is known as…….....…

(vi) ………… is the quantity of heat needed to raise the temperature of one mole of a substance by 1°C

(vii) $C_p - C_v =$ ....….

(viii) …......... $= \Delta H - T\Delta S$.

(ix) According to .... law of thermodynamics, $\Delta S_{total} = +$ ve.

(x) If $\Delta H = +$ve and $\Delta S = +$ve, the reaction is spontaneous at .... temperature

Ans: (i) Entropy (ii) spontaneous (iii) +ve, +ve (iv) –273°C
(v) Enthalpy (vi) molar heat capacity (vii) R (viii) ΔG
(ix) second (x) high

TRUE AND FALSE TYPE QUESTIONS

Write true or false for following statements:

(i) For every chemical reaction at equilibrium ΔG^0 is zero.

(ii) Entropy is not a state function because its value depends upon the condition of temperature and pressure.

(iii) During isothermal expansion of an ideal gas, there is no change in internal energy.

(iv) q and w are not state function but q+w is a state function.

(v) The enthalpy of neutralization of a strong acid by a strong base is always constant.

(v) For a spontaneous process $\Delta S_{system} = +$ve.

(vi) ΔH is not a state function.

(vii) The energy of universe is conserved while its entropy is increasing.

(viii) Volume is extensive property while temperature is intensive property.

(ix) At 0°C the entropy of a perfectly crystalline solid is zero.

(x) Hess' law is a corollary of the first law of thermodynamics.

Ans: (i) False (ii) False (iii) True (iv) True (v) True
(vi) False (vii) True (viii) True (ix) False (x) True

MATCH THE COLUMNS

I. Match the sign of ΔH in Column I with the sign of ΔS Column II and sign of ΔG in Column III for a Spontaneous & Non spontaneous reaction.

S.N.	Column I-[ΔH]	Column II-[ΔS]	Column III-[ΔG]
1.	–Ve	(a) +Ve	(p) +Ve at low temperature
2.	+Ve	(b) +Ve	(q) -Ve at High Temperature
3	+Ve	(c) +Ve	(r) –Ve at all temperature
4.	+Ve	(d) -Ve	(s) +Ve at all Temperature

II. Match the Process in Column I with the parameters in Column II and parameters in Column III

S.N.	Column I	Column II	Column III
1.	Isochoric Process	(a) Temperature constant	(p) w is not zero
2.	Isothermal Process	(b) Volume Constant	(q) T not constant
3.	Adiabatic Process	(c) Pressure Constant	(r) $\Delta U=0$
4.	Isobaric Process	(d) Heat Constant	(s) Work is zero

ANS :

MATCH-I : 1. b, r 2. a, q 3. d, s 4. c, p

MATCH-II : 1. b, s 2. a, r 3. d, q 4. c, p

ASSERTION AND REASON TYPE QUESTIONS

In the following questions a statement of assertion (A) followed by a statement of Reason (R) is given. Choose the correct option out of the choices given below for each question.

(i) A and R both are correct, and R is correct explanation of A.

(ii) A and R both are correct, but R is not the correct explanation of A.

(iii) A is true but R is false.

(iv) A and R both are false.

1. Assertion (A): Enthalpy of graphite is lower than that of diamond.
   Reason (R): Entropy of graphite is greater than that of diamond.
2. Assertion (A): Enthalpy of formation of $H_2O(l)$ is greater than that of $H_2O(g)$.
   Reason (R): Enthalpy change is negative for condensation reaction, $H_2O(g)$ $H_2O(l)$
3. Assertion (A): ΔH and ΔU are same for the reaction $N_2(g) + O_2(g) \longrightarrow 2NO(g)$
   Reason (R): All the reactants and products are gases.
4. Assertion (A): if both ΔH^0 and ΔS^0 are positive then the reaction will be spontaneous at high temperature
   Reason (R): All processes with positive entropy change are spontaneous.
5. Assertion (A): Enthalpy of formation of HCl is equal to bond energy of HCl.
   Reason (R): Enthalpy of formation and bond energy both involve the formation of one mole of HCl from the elements.

6. ASSERTION : The standard free energy change for all the spontaneously occurring reaction are negative

   REASON: The standard free energies of the elements in their standard states at 1 bar and 298 K are taken as zero

7. ASSERTION: Enthalpy and Entropy of any elementary substances in the standard state are taken as zero.

   REASON: At absolute zero, particles of perfectly crystalline substances become completely motionless.

8. ASSERTION : Enthalpy of Neutralization of CH_3COOH by NaOH is less than that of HCl by NaOH

   REASON : CH_3COOH is a weak acid hence need energy to ionize completely.

9. ASSERTION : The heat absorbed during the isothermal expansion of an ideal gas against vacuum is zero

   REASON: Internal energy of a gas depends upon Pressure of the gas.

10. ASSERTION : A reaction which is spontaneous and accompanied by decrease of randomness must be exothermic

    REASON : All exothermic reactions are accompanied by decrease of randomness.

Ans: 1. (ii) 2. (i) 3. (ii) 4. (iii) 5. (i) 6. (ii) 7. (iv) 8. (i) 9. (iii) 10. (iii)

ONE WORD ANSWER TYPE QUESTIONS

1. 'w' amount of work is done by the system and 'q' amount of heat is supplied to the system. What type of system would it be?
2. What is the work done in free expansion of an ideal gas?
3. What is the sign of ΔG^0 for spontaneous reaction?
4. Write the relation between ΔH and ΔU for $H_2(g) + I_2(g) \rightarrow 2HI(g)$.
5. Write the SI unit of entropy.
6. Name the calorimeter used to measureΔU.
7. What is the standard enthalpy of formation of graphite?
8. What is the sign of ΔH for $H_2(g) \longrightarrow 2H(g)$?
9. If $K_c = 1$, what will be the value of ΔG?
10. An exothermic reaction is spontaneous at all temperature. What is the sign of ΔS?

Ans: 1. Closed system 2. W = 0 3. ΔG = –ve 4. ΔH = ΔU 5. $J\,K^{-1}\,mol^{-1}$

6. Bomb calorimeter 7. Zero 8. ΔH = +ve 9. Zero 10. ΔS = +ve

1-MARK QUESTIONS

1. Name the thermodynamic system to which following belong:

   (i) Human body (ii) Milk in Thermos flask (iii) Tea in steel kettle

2. Identify State functions out of the following: Enthalpy, Entropy, Heat, Temperature, Work, Gibb's free energy.

3. Give two examples of state function.

4. Write the mathematical statement of first law of thermodynamics.

5. Predict the internal energy change for an isolated system?

   [**Ans.** Zero]

6. Why ΔH is more significant than ΔU?

7. Write one example each of extensive and intensive properties.

8. Write a chemical equation in which ΔH and ΔU are equal.

9 Write the relationship between ΔH and ΔU for the reaction:

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

10. Define standard enthalpy of formation.

11. Why is the standard enthalpy of formation of diamond not zero although it is an element?

12. The enthalpy of atomization of CH_4 is 1665 kJ mol^{-1}. What is the bond enthalpy of C – H bond?

    [**Ans.** 416.25 kJ]

13. Identify the species for which $\Delta_f H^\theta = O$, at 298 K; $O_3(g)$, $Br_2(g)$, $Cl_2(g)$ $CH_4(g)$.

14. For the reaction $2Cl(g) \longrightarrow Cl_2(g)$; what are the sign of ΔH and ΔS ?

15. Out of Diamond and Graphite which has higher entropy?

16. Write an example of endothermic spontaneous reaction.

17. State second law of thermodynamics.

18. State third law of thermodynamics.

19. Which has more entropy? 1 mol $H_2O(l)$ at 25°C or 1 mol $H_2O(l)$ at35°C.

20. At what temperature the entropy of a perfectly crystalline solid is zero?

21. For a certain reaction $\Delta G^0 = 0$,what is the value of K_c?

22. For a reaction both ΔH and ΔS are negative. Under what conditions does the reaction occur.

2-MARKS QUESTIONS

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

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

3. Classify the following as extensive or intensive properties :-
   Heat capacity, Density, Temperature, Molar heat capacity.

4. Derive the relationship between ΔH and ΔU.

5. Derive the relationship $C_p - C_v = R$.

6. State and explain Hess's Law of Constant Heat Summation with a suitable example.

7. Derive the relationship between ΔH and ΔU.
   Given, $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$; $\Delta_r H^\theta = -92.4$ kJ mol^{-1};
   What is the standard enthalpy of formation of NH_3 gas? [**Ans.** –46.2 kJ mol^{-1}]

8. Calculate the enthalpy change for the reaction:
   $H_2(g) + Br_2(g) \longrightarrow 2\,HBr(g)$.
   Given the bond enthalpies H_2, Br_2 and HBr are 435 kJ mol^{-1}, 192 kJ mol^{-1} and 368 kJ mol^{-1} respectively. [**Ans.** –109 kJ mol^{-1}]

9. Is the bond dissociation enthalpy of all the four C – H bonds in CH_4 same? Give reason in support of your answer.

10. Define the term entropy. Write its unit. How does entropy of a system change on increasing temperature?

11. The equilibrium constant for a reaction is 10.What will be the value of ΔG^θ? $R = 8.314$J K^{-1} mol^{-1}, T = 300K. [**Ans.** –5.744 kJ mol^{-1}]

12. The ΔH and ΔS for $2Ag_2O(s) \longrightarrow 4Ag(s) + O_2(g)$ are given + 61.17kJ mol^{-1} and +132 J K^{-1}mol^{-1} respectively. Above what temperature will the reaction be spontaneous? [**Ans.** >463.4 K]

3-MARKS QUESTIONS

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

2. Calculate the maximum work obtained when 0.75 mole of an ideal gas expands isothermally and reversibly at 27°C from a volume of 15L to 25L. [**Ans.** –955.7 J]

3. Calculate the number of kJ necessary to raise the temperature of 60 g of aluminium from 35 to 55°C . Molar heat capacity of Al is 24 J mol^{-1} K^{-1}. [**Ans.** 1.067 kJ]

4. The reaction of cyanamide, NH_2CN (s), with Dioxygen was carried out in a bomb calorimeter, and ΔU was found to be –742.7 kJ mol^{-1} at 298K. Calculate Enthalpy change for the reaction at 298K,

   $NH_2\ CN\ (s) + 3/2\ O_2\ (g) \longrightarrow N_2\ (g) + CO_2\ (g) + H_2O\ (l)$

   [**Ans.** –741.5 kJ mol^{-1}]

5. The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are –890.3 kJ mol^{-1}, –393.5 kJ mol^{-1} and –285.8 kJ mol^{-1} respectively. Calculate enthalpy of formation of methane gas. [**Ans.** –74.8 kJ mol^{-1}]

6. Explain the Born Haber Cycle to determine the lattice enthalpy of NaCl.

7. Enthalpies of formation of CO(g), CO_2(g), N_2O(g) and N_2O_4 (g) are –110, –393, 81 and 9.7 kJ mol-1 respectively. Find the value of Δ_rH for the reaction; $N_2O_4(g) + 3\ CO\ (g) \longrightarrow N_2O\ (g) + 3CO_2\ (g)$

   [**Ans.** –777.7 kJ mol^{-1}]

8. The combustion of 1 mol of benzene takes place at 298K .After combustionCO_2 and H_2O are formed and 3267 kJ mol^{-1} of heat is liberated. Calculate $\Delta_f H^0\ (C_6H_6)$.

   Given: $\Delta_fH^0\ (H_2O) = -286$ kJ mol^{-1}, $\Delta_fH^0(CO_2) = -393$ kJ mol^{-1}

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

9. Calculate the standard enthalpy of formation of $CH_3OH(l)$ from the following data:

$CH_3OH\,(l) + 3/2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(l);\ \Delta_cH^\theta = -726\ \text{kJ mol}^{-1}$

$C\,(g) + O_2\,(g) \longrightarrow CO_2\,(g);\ \Delta_f\,H^\theta = -393\ \text{kJ mol}^{-1}$

$H_2(g) + 1/2\,O_2\,(g) \longrightarrow H_2O\,(l);\ \Delta_f\,H^\ominus - 286\ \text{kJ mol}^{-1}$

[**Ans.** –239 kJ mol^{-1}]

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

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

[**Ans.** T > 2000 K]

12. Reaction $X(s) \rightarrow Y(g)$ ΔH = +ve is spontaneous at temperature 'T'. Determine
    (i) Sign of ΔS for this reaction.
    (ii) Sign of ΔG for $Y \longrightarrow X$
    (iii) Sign of ΔG at a temperature < T

5-MARKS QUESTIONS

1. (a) What is reversible process in Thermodynamics?
   (b) Name the thermodynamic processes for which : (i) Δq = 0 (ii) ΔU = 0 (iii) ΔV = 0 (iv) ΔP = 0
   (c) Water decomposes by absorbing 286.2 kJ of electrical energy per mole. When H_2 and O_2 combine to form one mole of H_2O, 286.2 kJ of heat is produced. Which thermodynamic law is proved? Write its statement.

2. (a) Why work done in isothermal reversible process is always maximum?

[Hint: $q_v = \Delta U$ and $q_p = \Delta H$]

   (b) It has been found that 221.4 J is needed to heat 30g of ethanol from 15°C to 18°C. Calculate (a) specific heat capacity, and (b) molar heat capacity of ethanol.

[**Ans.** (a) 2.46 $Jg^{-1}\,°C^{-1}$, (b) 113.2 J $mol^{-1}\,°C^{-1}$]

3. (a) Differentiate the terms Bond dissociation enthalpy & Bond Enthalpy.

(b) Calculate enthalpy change for the process $CCl_4(g) \longrightarrow C(g) + 4Cl(g)$ and calculate Bond enthalpy of C–Cl bond in CCl_4.

Given: $\Delta_{vap} H^{\theta} = 30.5$ kJ mol^{-1}; $\Delta_f H^{\theta}(CCl_4) = -135.5$ kJ mol^{-1};

$\Delta_a H^{\theta}(C) = 715$ kJ mol^{-1} and $\Delta_a H^{\theta}(Cl_2) = 242$ kJ mol^{-1}

[**Ans.** 1304 kJ mol^{-1}, 326 kJ mol^{-1}]

4. Predict the sign of ΔS for the following changes:

(i) Freezing of water.

(ii) $C(graphite) \longrightarrow C(diamond)$

(iii) $H_2(g)$ at 298 k and 1 bar $\longrightarrow$ $H_2(g)$ at 298 k and 10 bar

(iv) $H_2(g) + I_2(g) \longrightarrow 2HI(g)$

(v) $2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$

5. (i) Define Gibbs Energy. Give its mathematical expression. What is Gibb's energy criteria of Spontaneity.

(ii) For the reaction:

$2A(g) + B(g) \rightarrow 2D(g)$, $\Delta U^{\theta} = -10.5$ kJ and $\Delta S^{\theta} = -44.1$ J K^{-1}.

Calculate ΔrG° for the reaction, and predict whether the reaction will occur spontaneously.

[**Ans.** $\Delta_r G^{\theta} = +0.16$ kJ, Non-spontaneous]

HOTS QUESTIONS

1. Does entropy increase or decrease when egg is boiled?

Ans.: On boiling egg, entropy decreases as due to denaturation, the helical structure of protein become more complicated and random coiled structure.

2. 10 g of argon is compressed isothermally and reversibly at a temperature of 27°C from 10 L to 5 L. Calculate q, w, ΔU and ,ΔH.

Solution: $q = -2.303\ nRT \log V_2 / V_1 = -2.303 \times 10/40$ mol $\times$ 2 Cal K^{-1} mol^{-1} $\times$ 300 K $\times \log 5/10 = -103.635$ Cal

For isothermal compression ΔU = 0

$W = \Delta U - q = 0 - (-103.635) = +103.635$ Cal

Also when temperature is constant,

$$PV = \text{constant}, \quad \Delta H = \Delta U + \Delta(PV) = 0 + 0 = 0$$

3. 1 mole of an ideal gas expand isothermally and reversibly from a pressure of 10 atm to 1 atm at 300 K. Calculate the height to which an object of 50 kg can be lifted by this expansion.

Solution: $w_{(exp.)} = -[2.303\ nRT \log (P_i / P_f)]$

$$= -2.303 \times 1\ \text{mol} \times 8.314\ \text{J K}^{-1}\text{mol}^{-1} \times 300\ \text{K} \times \log 10/1$$

$$= 5.74 \times 10^3\ \text{J}$$

Now, $mgh = 5.74 \times 10^3$ J or $50\ \text{kg} \times 9.81\ \text{m s}^{-2} \times h = 5.744 \times 10^3$ J

$\therefore\ h = 11.7$ m

UNIT TEST-I

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. For the reaction $2Cl(g) \longrightarrow Cl_2(g)$; what are the sign of ΔH and ΔS? [1]
2. Write an example of endothermic spontaneous reaction. [1]
3. 'w' amount of work is done by the system and 'q' amount of heat is supplied to the system. What type of system would it be? [1]
4. In a process 701 J of heat is absorbed by a system and 394 J work is done by the system. What is the change in internal energy for the process? [2]
5. State and explain Hess's Law of Constant Heat Summation with a suitable example. [2]
6. Calculate the number of kJ necessary to raise the temperature of 60 g of aluminium from 35 to 55°C . Molar heat capacity of Al is $24\ J\ mol^{-1}\ K^{-1}$. [3]
7. Calculate the standard enthalpy of formation of $CH_3OH(l)$ from the following data:

   $CH_3OH(l) + 3/2\ O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$;
   $\Delta_c H^\theta = -726\ kJ\ mol^{-1}$

   $C(g) + O_2(g) \longrightarrow CO_2(g)$; $\Delta_f H^\theta = -393\ kJ\ mol^{-1}$

   $H_2(g) + \frac{1}{2} O_2(g) \longrightarrow H_2O(l)$; $\Delta_f H^\theta = -286\ kJ\ mol^{-1}$

8. (a) For oxidation of iron, $4\,Fe(s) + 3\,O_2(g) \longrightarrow 2\,Fe_2O_3(s)$ entropy change is $-549.4\ J\ K^{-1}\ mol^{-1}$ at 298 K. In spite of negative entropy change of this reaction, why is the reaction spontaneous? ($\Delta r\,H^\circ$ for this reaction is $-1648\ kJ\ mol^{-1}$) [2]

   (b) For the reaction: $2A(g) + B(g) \longrightarrow 2D(g)$, $\Delta U^\circ = -10.5\ kJ$ and $\Delta S^\circ = -44.1\ J\ K^{-1}$. Calculate ΔG^0 for the reaction, and predict whether the reaction will occur spontaneously. [3]

UNIT TEST-II

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

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1. Among following, the property which is intensive is [1]

   (a) Mass (b) Volume

   (c) Temperature (d) Length

2. Heat of formation is zero for which substance. [1]

   (a) $CaCO_3$ (b) HCl

   (c) Carbon (Diamond) (d) Carbon (Graphite)

3. What is an isochoric process and what is the work value associated with isochoric process? [1]

4. Name the law which is based on the statement "enthalpy is a state function." [1]

5. Heat of combustion of ethane is -1560 KJ mol^{-1} while that of acetylene is -1296KJ mol^{-1}. Which is better fuel and why? [1]

6. Derive the expression for work done in isothermal reversible expansion. [3]

7. Define standard enthalpy of combustion and standard enthalpy of formation taking C_2H_6 (gas) molecule in both cases. [2]

8. State 'Hess law of Constant heat summation'. The molar heat of combustion of $C_2H_2(g)$, C(Graphite) and $H_2(g)$ are 310.62kcal, 94.05 kcal and 68.32 kcal respectively. Calculate the standard heat of formation of $C_2H_2(g)$. [3]

9. The reaction of Cyanamid, $NH_2CN(s)$, with dioxygen was carried out in a bomb calorimeter, and ΔU was found to be -742.7 $kJmol^{-1}$ at 298 K. Calculate enthalpy change for the reaction at 298 K.

$NH_2CN(s) + 1.5\ O_2(g) \rightarrow N_2(g) + CO_2(g) + H_2O(l)$ [3]

10. (i) State the First Law of Thermodynamics.

(ii) Define Entropy. What is the effect of temperature on Entropy?

(iii) Two moles of an ideal gas are held by a piston under 5 atm pressure at 300 K. The pressure is suddenly released to 0.4 atm and the gas is allowed to expand isothermally. Calculate W, q, ΔE, ΔH. [5]

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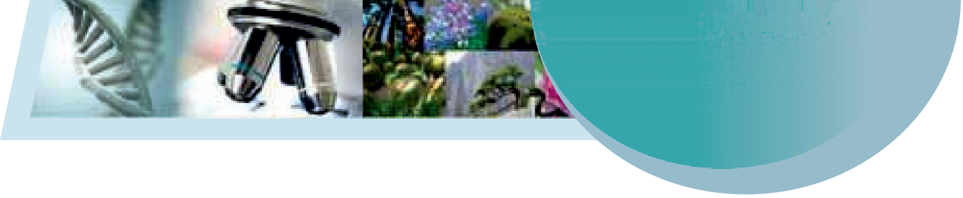
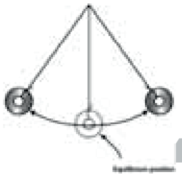




Chapter - 6

Equilibrium

FAST TRACK : QUICK REVISION

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

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

$$aA + bB \rightleftharpoons cC + dD$$

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

- **Relationship between K_p and K_c :**

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

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

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

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

- **Predicting the direction of reaction :**

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

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

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

- **Ostwald's dilution law :** Degree of dissociation of weak electrolyte, $\alpha = \sqrt{\frac{K}{C}}$

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

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

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

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

- **Conjugate Acid or Base : Acid-base pair which differ by H^+ ion.**

  Species – H^+ = Conjugate base

  Species + H^+ = Conjugate acid

- **pH of solution :**

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

  $pH + pOH = pK_w = 14$ at 298K

- **Common ion effect** : The depression of ionisation of weak electrolyte by the presence of common ion from a strong electrolyte is called common ion effect. For example degree of dissociation of NH_4OH decreases in the presence of strong electrolyte NH_4Cl.

- **Hydrolysis of salts and pH of their solutions :** Hydrolysis of salt is defined as the reaction of cation or anion with water as a result of which the pH of water changes.

1. Salts of strong and strong bases (*e.g.*, NaCl) do not hydrolyse. The solution pH will be 7.

2. Salts of weak acids and strong bases (*e.g.*, CH_3COONa) hydrolyse, pH >7 (The anion acts as a base).

$$\underset{}{X^-} + H_2O \rightleftharpoons \underset{\text{(Weak acid)}}{HX} + \underset{\text{(Weak base)}}{OH^-}$$

$$pH = 7 + \frac{1}{2}(pK_a + \log C)$$

3. Salt of strong acids and weak bases (*e.g.*, NH_4Cl) hydrolyse, pH < 7. (The cation acts as an acid).

$$M^+ + H_2O \rightleftharpoons MOH + H^+$$

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

4. Salt of weak acids and weak base (*e.g.*, CH_3COONH_4) hydrolyse. The cation acts as an acid and anion as a base but whether the solution is acidic or basic depends upon the relative values of K_a and K_b for these **ions.**

$$M^+ + X^- + H_2O \rightleftharpoons MOH + HX$$

$$pH = 7 + \frac{1}{2}(pK_a - pK_b)$$

- **Buffer solutions :** The solutions, which resist the change in pH on dilution or addition of small amounts of acid or base, are called buffer solutions.

- **Basic buffer :** Solution of weak base and its salt with strong acid, For *e.g.*, $NH_4OH + NH_4Cl$

- **Acidic buffer :** Solution of weak acid and its salt with strong base, For *e.g.*, $CH_3COOH + CH_3COONa$.

- **Henderson Hasselbalch Equation for the pH of Buffer solution—**

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

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

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

For $A_xB_y \overset{aq}{\rightleftharpoons} xA^{y+} + yB^{x-}$

$$K_{sp} = [A^{y+}]^x [B^{x-}]^y = (xs)^x (ys)^y = x^x . y^x . s^{(x+y)}$$

where s = Molar solubility

If ionic product < K_{sp} ; salt remain dissolve.

If ionic product > K_{sp} ; salt will be precipitated.

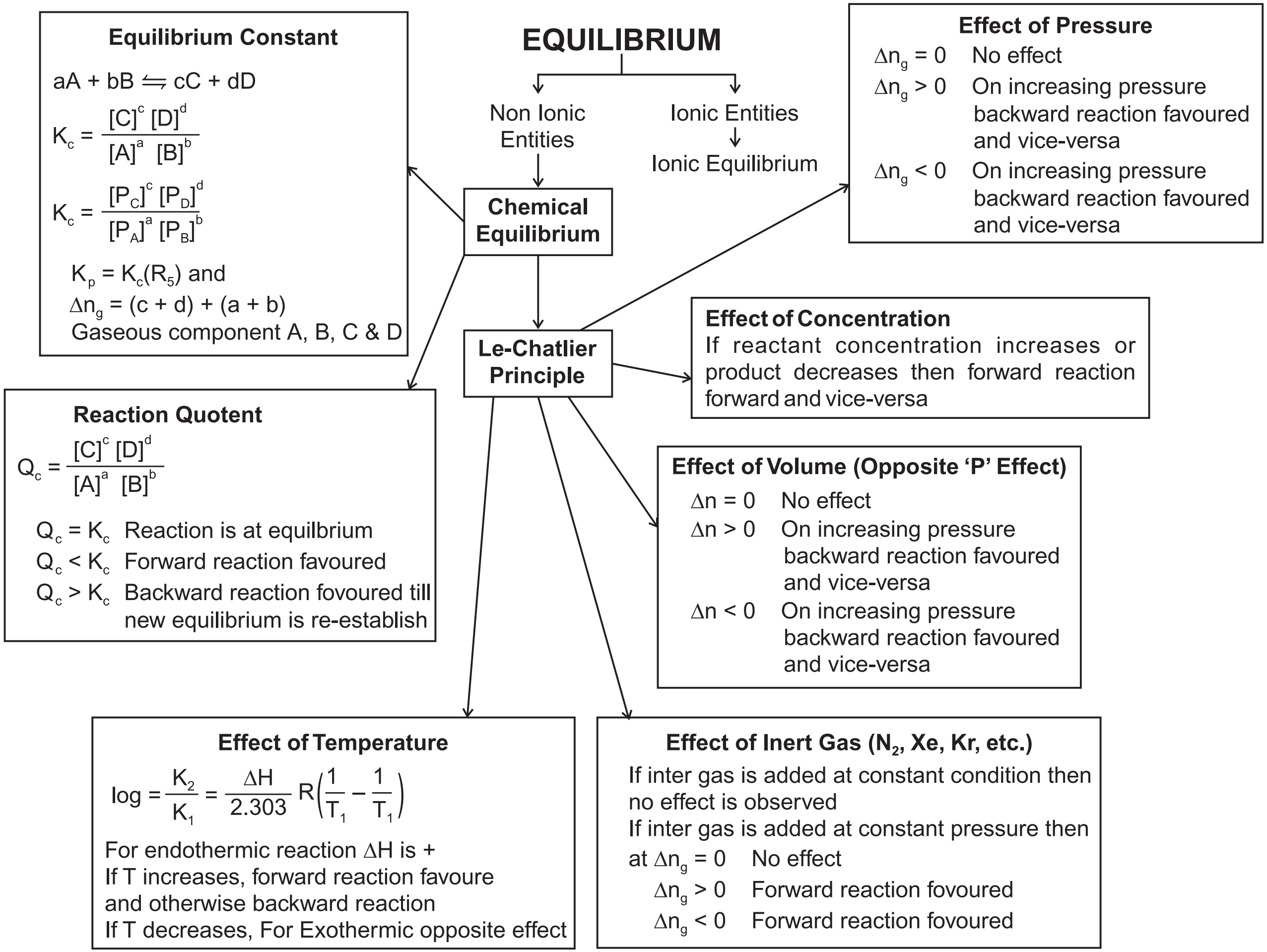
- **Relationship between solubility (s) and solubility product (K_{sp}).**

$$K_{sp} = x^x . y^y . s^{x+y}$$

For binary salts (*e.g.*, AgCl, AgBr, AgI) $\qquad K_{sp} = s^2$

For Ternary salts (*e.g.*, PbI_2) $\qquad K_{sp} = 4s^3$

MIND MAP : EQUILIBRIUM



EQUILIBRIUM

Non Ionic Entities → Chemical Equilibrium → Le-Chatlier Principle

Ionic Entities → Ionic Equilibrium

Equilibrium Constant

$aA + bB \leftrightharpoons cC + dD$

$$K_c = \frac{[C]^c\ [D]^d}{[A]^a\ \ [B]^b}$$

$$K_c = \frac{[P_C]^c\ [P_D]^d}{[P_A]^a\ [P_B]^b}$$

$K_p = K_c(R_5)$ and

$\Delta n_g = (c + d) + (a + b)$

Gaseous component A, B, C & D

Reaction Quotent

$$Q_c = \frac{[C]^c\ [D]^d}{[A]^a\ \ [B]^b}$$

$Q_c = K_c$ Reaction is at equlibrium

$Q_c < K_c$ Forward reaction favoured

$Q_c > K_c$ Backward reaction fovoured till new equilibrium is re-establish

Effect of Pressure

$\Delta n_g = 0$ No effect

$\Delta n_g > 0$ On increasing pressure backward reaction favoured and vice-versa

$\Delta n_g < 0$ On increasing pressure backward reaction favoured and vice-versa

Effect of Concentration

If reactant concentration increases or product decreases then forward reaction forward and vice-versa

Effect of Volume (Opposite 'P' Effect)

$\Delta n = 0$ No effect

$\Delta n > 0$ On increasing pressure backward reaction favoured and vice-versa

$\Delta n < 0$ On increasing pressure backward reaction favoured and vice-versa

Effect of Temperature

$$\log = \frac{K_2}{K_1} = \frac{\Delta H}{2.303} R\left(\frac{1}{T_1} - \frac{1}{T_1}\right)$$

For endothermic reaction ΔH is +

If T increases, forward reaction favoure and otherwise backward reaction

If T decreases, For Exothermic opposite effect

Effect of Inert Gas (N_2, Xe, Kr, etc.)

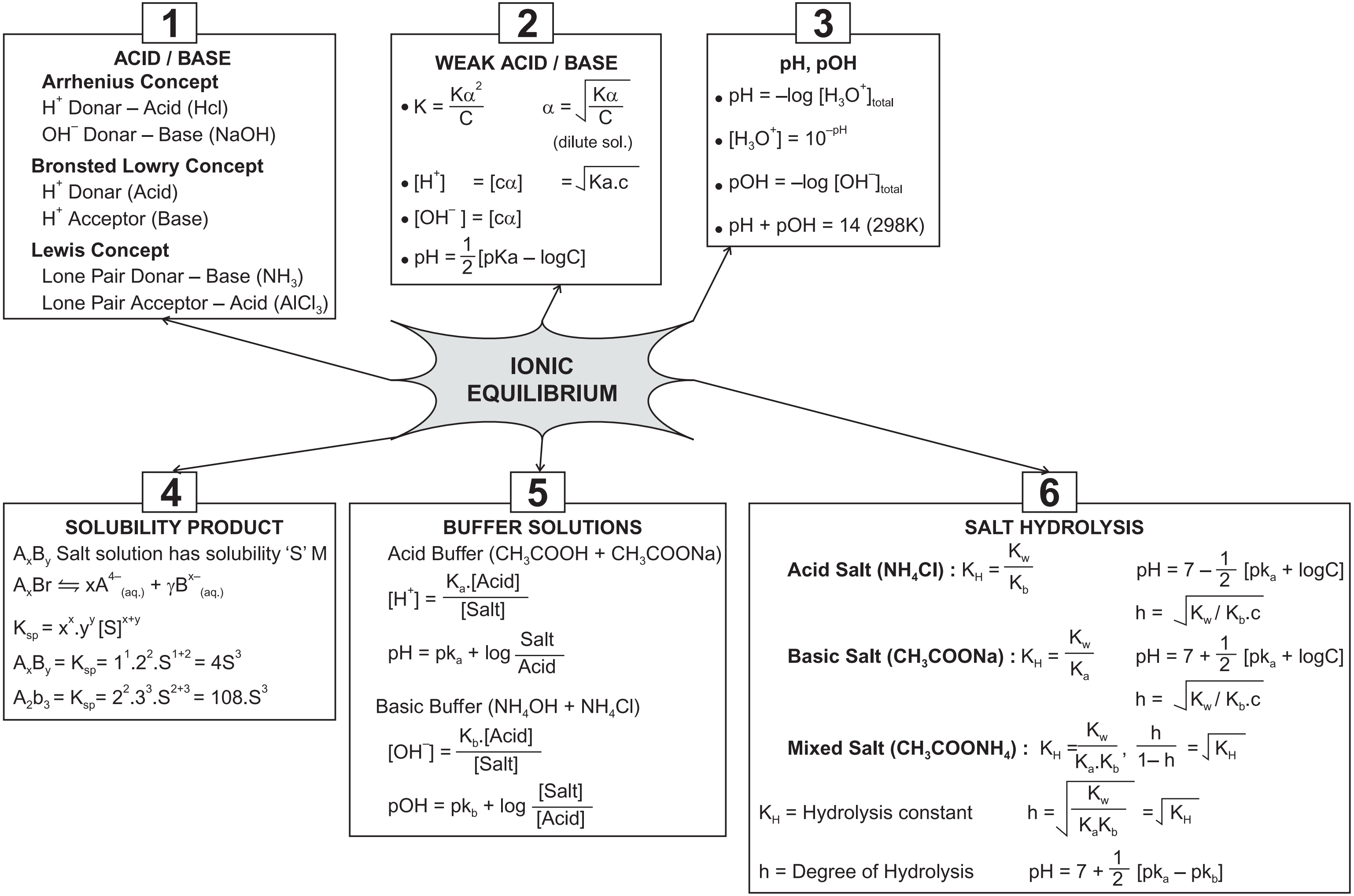
If inter gas is added at constant condition then no effect is observed

If inter gas is added at constant pressure then

at $\Delta n_g = 0$ No effect

$\Delta n_g > 0$ Forward reaction fovoured

$\Delta n_g < 0$ Forward reaction fovoured



IONIC EQUILIBRIUM

1

ACID / BASE

Arrhenius Concept

H^+ Donar – Acid (Hcl)

OH^- Donar – Base (NaOH)

Bronsted Lowry Concept

H^+ Donar (Acid)

H^+ Acceptor (Base)

Lewis Concept

Lone Pair Donar – Base (NH_3)

Lone Pair Acceptor – Acid ($AlCl_3$)

2

WEAK ACID / BASE

- $K = \frac{K\alpha^2}{C}$ $\quad \alpha = \sqrt{\frac{K\alpha}{C}}$ (dilute sol.)
- $[H^+] = [c\alpha] = \sqrt{Ka.c}$
- $[OH^-] = [c\alpha]$
- $pH = \frac{1}{2}[pKa - \log C]$

3

pH, pOH

- $pH = -\log [H_3O^+]_{total}$
- $[H_3O^+] = 10^{-pH}$
- $pOH = -\log [OH^-]_{total}$
- $pH + pOH = 14$ (298K)

4

SOLUBILITY PRODUCT

A_xB_y Salt solution has solubility 'S' M

$A_xBr \leftrightharpoons xA^{4-}_{(aq.)} + \gamma B^{x-}_{(aq.)}$

$K_{sp} = x^x.y^y [S]^{x+y}$

$A_xB_y = K_{sp} = 1^1.2^2.S^{1+2} = 4S^3$

$A_2b_3 = K_{sp} = 2^2.3^3.S^{2+3} = 108.S^3$

5

BUFFER SOLUTIONS

Acid Buffer ($CH_3COOH + CH_3COONa$)

$$[H^+] = \frac{K_a.[Acid]}{[Salt]}$$

$$pH = pk_a + \log \frac{Salt}{Acid}$$

Basic Buffer ($NH_4OH + NH_4Cl$)

$$[OH^-] = \frac{K_b.[Acid]}{[Salt]}$$

$$pOH = pk_b + \log \frac{[Salt]}{[Acid]}$$

6

SALT HYDROLYSIS

Acid Salt (NH_4Cl) : $K_H = \frac{K_w}{K_b}$ $\quad pH = 7 - \frac{1}{2}[pk_a + \log C]$

$h = \sqrt{K_w / K_b.c}$

Basic Salt (CH_3COONa) : $K_H = \frac{K_w}{K_a}$ $\quad pH = 7 + \frac{1}{2}[pk_a + \log C]$

$h = \sqrt{K_w / K_b.c}$

Mixed Salt (CH_3COONH_4) : $K_H = \frac{K_w}{K_a.K_b}, \frac{h}{1-h} = \sqrt{K_H}$

K_H = Hydrolysis constant $\quad h = \sqrt{\frac{K_w}{K_aK_b}} = \sqrt{K_H}$

h = Degree of Hydrolysis $\quad pH = 7 + \frac{1}{2}[pk_a - pk_b]$

CASE BASED STUDY QUESTIONS

PASSAGE -1

According to Arrehenius theory, acids are substances that dissociates in water to give hydrogen ions and bases are substances that produce hydroxyl ions.

Bronsted-Lowry gave a more general definition of acids and bases. According to Bronsted-Lowry theory , acid is a substance that is capable of donating a hydrogen ion and bases are substance which are capable of accepting of a hydrogen ion.

According to Lewis Acids and Bases , an acid is a species which accepts electron pair and base which donates an electron pair

Answer the following questions

1. Which of the following acts as both Bronsted acid as well as Bronsted base

   (A) HCO_3 (B) O_2

   (C) HS^- (D) HPO_3^{2-}

2. Arrange the following in order of increasing basic strength

   OH^-; CH_3COO^-, Cl^-, $C_2H_5O^-$

3. Define lewis acid, Give example.

4. All Arrhenius acid are Bronsted acid but Arrhenius bases are not Bronsted base. Explain.

   OR

   Classify the following as lewis acid and lewis base and show hot they acts as such

   (A) $AlCl_3$ (B) H_2O

ANS. 1. HCO_3^-, HS^- 2. $Cl^- < CH_3COO^- < OH^- < C_2H_5O^-$ 3. Definition

4. all Arrhenius acid gives H^+ in aqueous solution while Bronsted acids also gives H^+ and hence all Arrhenius acids are Bronsted acids but Arrhenius base given OH^- in aqueous solution while Bronsted bases accept H^+.

OR

$AlCl_3$: lewis acid as it accept electron pair

H_2O : Lewis base as it can donate lone pair of electron present on oxygen atom.

PASSAGE -2

Common Ion effect is the phenomenon in which weak acid or weak base dissociation is suppressed due to the presence of the common ion provided by the strong electrolyte . For example the dissociation of CH_3COOH is suppressed by the addition of the CH_3COONa. Similarly the dissociation of NH_4OH is suppressed due to presence of NH_4Cl.This occurs due to the Le-Chatlier's principle .

Due to common ion effect there is decrease in solubility of an ionic precipitate by addition to the solution of a soluble compound with an ion common with then precipitate. In water treatment process highly soluble sodium carbonate salt is added to precipitate out sparingly soluble salt calcium carbonate. Salting out process used in manufacturing soap is also benefit from the common ion effect. Common ion effect also help in maintaining pH of a buffer solution.

Answer the following questions:

1. What is common ion effect?
2. Why salt is added to soap during soap formation.
3. Name the principle common ion is based on.
4. Both cation of 2th and 4th group are precipitated as sulphides, but 4th group radicals do no get precipitated when H_2S is passed through solution for 2nd group why.

Or

3rd group radicals are precipitated as hydroxides by addition of NH_4OH in presence of NH_4Cl, what is the role of NH_4Cl.

MULTIPLE CHOICE QUESTION (MCQ)

1. For the hypothetical reactions, the equilibrium constant (k) values are given

   $A \rightleftharpoons B : k_1 = 2$

   $B \rightleftharpoons C : K_2 = 4$

   $C \rightleftharpoons D : K_3 = 8$

   The equilibrium constant (K) for the reaction $A \rightleftharpoons D$ is

   (a) 48 (b) 24 (c) 12 (d) 64

2. The equilibrium constant for the reaction

   $SO_2(g) + \frac{1}{2} O_2(g) \rightleftharpoons SO_3(g)$ is 5×10^{-2} atm$^{-1/2}$

   The equilibrium constant for the reaction

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

   (a) 100 atm (b) 25×10^{-4} atm (c) 400 atm (d) 125×19^{-6} atm$^{-3/2}$

3. $A(g) + 3B(g) \rightleftharpoons 4C(g)$ initial concentration of A is equal to that of B. The equilibrium concentrations of A and C are equal. What is the equilibrium constant for

   $4C(g) \rightleftharpoons A(g) + 3B(g)$

   (a) 4 (b) 1/8 (c) 64 (d) 16

4. The equilibrium reaction that is not affected by volume change at constant temperature is

   (a) $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ (b) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   (c) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ (d) $H_2(l) + CO_2(g) \rightleftharpoons H_2CO_3(l)$

5. For the reaction $CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$, the value of K_c/K_p is equal to

   (a) RT (b) $(RT)^2$ (c) 1/RT (d) 1.0

6. At 90°C pure water has $K_w = 10^{-12}$. The solution with pH value 6.5 is

   (a) Acidic (b) Basic (c) Amphoteric (d) Data insufficient

7. 40 ml of 0.1 M NH_4OH is mixed with 20 mL of 0.1 M HCl. What is the pH of the mixture? (pK_b of ammonia solution = 4.74)

   (a) 4.74 (b) 2.26 (c) 9.26 (d) 5

8. Identify Bronsted Lowry Acids in the reaction

$$\underset{(X)}{[Al(H_2O)_6]^{3+}} + \underset{(Y)}{HCO_3^-} \rightleftharpoons \underset{(P)}{[Al(H_2O)_5(OH)]^{2+}} + \underset{(Q)}{H_2CO_3}$$

   (a) X, Y (b) Y, P (c) P, Q (d) X, Q

9. The pK_a of weak acid HA is 4.80 and pK_b of a weak base BOH is 4.78. The pH of an aqueous solutions of corresponding salt BA will be

   (a) 7.01 (b) 4.79 (c) 9.22 (d) 10.0

10 If 'p' M is the solubility of $KAl(SO_4)_2$, then K_{sp} is equal to

   (a) p^3 (b) $4p^4$ (c) p^4 (d) $4p^3$

TRUE AND FALSE TYPE QUESTIONS

1. Equilibrium state can be achieved if a reversible reaction is carried out in closed or open container.

2. For a reaction $2A(g) \rightleftharpoons B(g)$ $Q_c > K$ if 'A' is added maintaining $Q_c > K$, the reaction will move in backward direction.

3. For the reaction at equilibrium

   $CaCO_3 \rightleftharpoons CaO(s) + CO_2(g)$

   When CaO(s) is removed reaction moves in forward direction.

4. For a reaction $aA + bB \rightleftharpoons cC + dD$ at equilibrium $\Delta G^0 = 0$ always.

5. For a reaction at equilibrium $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$

   K = 4, the value of $\dfrac{K_b[HCl]^2}{K_f[H_2][Cl_2]}$ is 1.

6. For the electrolyte A_2B if K_{sp} is solubility product then its solubling 'S' M is $[K_{sp}]^{1/3} \div 4$.

7. HCO^-_3 is conjugate base of H_2CO_3.

8. H_2O can act as acid as well as base.

9. The pH of buffer solution remain same when any amount of dilution is done.

10. For a salt $AB_2(s)$ solution if Ionic product (I.P) > K_{sp}, then precipitation will take place.

Ans. 1. False	2. True	3. True	4. True	5. True
6. True	7. True	8. True	9. False	10. True

FILL IN THE BLANKS

1. At equilibrium rate of forward reaction is always equal to ....
2. K_p & K_c are .... for reaction at equilibrium of type $H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$.
3. If K_c for reaction $CH_3COOH(l) + C_2H_5OH \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$ is 4. Then Q_c and K_c are .... at equilibrium.
4. If A+B –70J/mol $\rightleftharpoons$ D, reaction temperature is increased then reaction moves in .... direction.
5. Presence of catalyst will ...... the equilibrium constant.
6. The conjugate acid of H_2O is ....
7. On dilution, the degree of dissociation of acetic acid will ....
8. The presence of NH_4Cl in NH_4OH solution will .... the degree of dissociation of NH_4OH.
9. If Ionic product (IP) < K_{sp} for a salt solution of AB, then addition of AB further .... lead to precipitation initially.
10. K_p is always equal to K_c if Δn_g is ....

Ans. 1. rate of backward reaction, 2. equal, 3. equal, 4. backward direction forward, 5. not change, 6. H_3O^+, 7. increase, 8. decrease, 9. will not, 10. zero.

MATCH THE COLUMNS

I. Match the reaction in Column I with the parameters in Column II and unit (M=Molarity) of K eq in Column III

S.N.	Column I	Column II	Column III
1.	$N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$, ΔH= -ve	(a) T increase then K increase	(p) M^0
2.	$2N_2(g)+2O_2(g) \rightleftharpoons 4NO(g)$, ΔH= +ve	(b) T increase then K decrease	(q) M^{2-}
3.	$2X(g) \rightleftharpoons Y(g)$, ΔH= +Ve	(c) P has not effect	(r) M^{-1}
4.	$PCl_5(g) \rightleftharpoons PCl_3(g)+Cl_2(g)$, ΔH= +ve	(d) Equilibrium move forward On addition of Xe Gas	(s) M

II. Match the parameter in Column I with the pH expression in Column II and examples in Column III

S.N.	Column I	Column II-pH	Column III
1.	Salt of weak acid and weak base	(a)7+ 0.5(pka +logC)	(p) NH_4Cl
2.	Salt of weak acid and strong base	(b) 7+ 0.5(pka -pkb)	(q) NaCl
3	Salt of strong acid and strong base	(c)7- 0.5(pkb +logC)	(r) CH_3COONa
4.	Salt of strong acid and weak base	(d) 0.5 (pkw)	(s) CH_3COONH_4

ANS. :

MATCH-I : 1. b, q 2. c, p 3. a, r 4. d, s

MATCH-II : 1. b, s 2. a, r 3. d, q 4. c, p

ASSERTION - REASON TYPE QUESTION

Each question contains statement-1 (assertion) and Statement-2 (Reason) Examine the statements carefully and mark the correct answer according to the instruction given below :

A. If both the statements are true and statement -2 is the correct explanation of statement-I

B. If both the statements are true but statement-2 is not the correct explanation of statement-I

C. If statement-I is true and statement-2 is false

D. If statement-I is false and statement-2 is true.

1. Statement-1 : The endothermic reactions are favoured at lower temperature and the exothermic reactions are favoured at higher temperature.
   Statement-2 : when a system in equilibrium is disturbed by changing the temperature, it will tend to adjust itself so as to overcome the effect of change.
2. Statement-1 : The melting point of ice decreases with increase of pressure
   Statement-2 : Ice contracts on melting.
3. Statement -1 : The gas phase reaction $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$ shifts to the right on increasing pressure.
   Statement-2 : When pressure increase, equilibrium shifts towards more number of moles.

4. Statement-1 : The chemical equilibrium is not static but dynamic in nature.
   Statement-2: The chemical equilibrium is a state in which two opposing process are proceeding at the same rate.

5. Statement-1 : The catalyst does not change the equilibrium constant.
   Statement-2 : Because for the catalysed reaction and uncatalysed reaction ΔH remains same and equilibrium constant depends on ΔH.

6. Statement-1 : If water is heated to 59°C, the pH will increase.
   Statement-2 : K_w increases with increase in temperature.

7. Statement-1: Addition of HCl(aq.) to CH_3COOH (aq.) decrease the ionisation of CH_3COOH (aq.).
   Statement-2 : Due to common ion effect H^+, ionisation of CH_3COOH decreases.

8. Statement-1: Sparingly soluble salts AB and XY_2 with the same solubility product, will have different solubility.
   Statement 2: Solubility of sparingly soluble salts depends upon solubility product.

9. Statement-1 : The ionisation constants of weak diprotic acid are in the order of $Ka_1 > Ka_2$.
   Statement-2 : Removal of H^+ from anion is difficult as compared to neutral atom.

10. Statement-1 : In a titration of weak acid with strong base, the pH at the half equivalence point is pK_a.
    Statement-2 : At half equivalence point, it will form acidic buffer at its maximum capacity where [Acid] = [Salt].

Ans.: 1. D, 2. A, 3. C, 4. A, 5. A, 6. D, 7. A, 8. B, 9. A, 10. A

ONE WORD ANSWER TYPE QUESTIONS

1. What is sum of pH + pOH at 25°C? [**Ans.** 14]
2. Write the Henderson Hasselbalch equation for acidic buffer
   Ans. $pH = pka + l0g \frac{[SALT]}{[ACID]}$
3. How is degree of dissociation related with concentration terms and Ka, for weak electrolyte. **Ans.** $\alpha = \sqrt{Ka / c}$
4. How NH_3 is defined as Lewis base?
   [**Ans.** It contain Lone paid of electrons]
5. How are K_p and K_c related? [**Ans.** $K_p = K_c (RT)^{\Delta n}$]
6. How does K affected for endothermic reaction if temperature is increased?
   [**Ans.** K get decreased]

7. What is the effect of catalyst on K? [**Ans.** K remains unaffected]

8. How is pH scale affected by increasing temperature? [**Ans.** pH scale gets contracted]

9. What is the conjugate base of HCO_3^-? [**Ans.** CO_3^{2-}]

10. What is the nature of CH_3COOH in conc. HCl solution? [**Ans.** Bronsted Base]

1-MARK QUESTIONS

1. Define physical equilibrium. Give an example also.

2. Fizz is observed when soda water bottle is opened. Why ?

3. Justify the statement : 'Both physical and chemical equilibrium are dynamic in nature'

4. State Law of Chemical equilibrium.

5. In a reversible reaction, the two substances are in equilibrium. If the concentration of each one is reduced to half, then what is the effect on the equilibrium constant ?

6. K_1 and K_2 are equilibrium constant for reactions (1) and (2)

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

(ii) $NO(g) \rightleftharpoons 1/2\,N_2(g) + 1/2\,O_2(g)$

Calculate the relation between K_1 and K_2.

7. Write the equilibrium constant expression for the following reaction :

$3\,Fe(s) + 4\,H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4\,H_2(g)$

8. Classify the equilibrium as homogeneous or heterogeneous :

$CH_3COOC_2H_5(aq.) + H_2O(l) \rightleftharpoons CH_3COOH(aq.) + C_2H_5OH\,(aq.)$

9. $K_p = \dfrac{(P_{NH_3})}{(P_{H_2})^{3/2}(P_{N_2})^{1/2}}$

Write the balanced chemical equation corresponding to the above expression.

10. Give the direction in which the reaction would proceed if $Q_c > K_c$.

11. $Hb(s) + O_2(g) \rightleftharpoons HbO_2(l)$

Predict the direction in which equilibrium gets shifted if partial pressure of $O_2(g)$ is lowered.

12. Discuss the position of equilibrium if the following reaction is carried out in the presence of catalyst.

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

13. Which of the following are Lewis acids ?

H_2O, BF_3, H^+, NH_4^+

14. Write the conjugate acids for the following Bronsted bases.

C_6H_5OH, H_2O

15. Write the conjugate bases for the following Bronsted acids.

H_2O, CH_3COOH.

16. Which of the following are Lewis acids ?

(a) H_2O (b) $AlCl_3$ (c) NH_4^+

17. Define Ostwald's dilution law.

18. SO_3^{2-} is Bronsted base or acid and why ?

19. Why pH of our blood remains almost constant at 7.4 though we quite often eat spicy food ?

20. pH of black coffee is 5.0 at 25ºC. Is black coffee acidic or basic ? [**Ans.** Acidic]

21. What will be the value of $(pK_a + pK_b)$ at 25ºC.

22. What will be the pH of 1 M KNO_3 solutions at 25ºC?

23. $CaCl_2(s) + H_2O(l) \rightleftharpoons CaCl_2(aq.) + Heat$

Discuss the solubility of $CaCl_2$ if temperature is increased.

24. Why does the solubility of CO_2 decrease with rise in temperature ?

25. The solubility of A_2X_3 is y mol dm^{-3}. Calculate its solubility product.

26. Write the K_{sp} expression for $Al(OH)_3$.

27. What is the condition for precipitation of a salt ?

28. Pridict whether the solution is acidic, basic or natural when NH_4NO_3 undergo hydrolysis.

29. Explain why pure NaCl precipitates out when HCl gas is passed through the solution of NaCl ?

30. Give the Henderson's Hasselbalch equation for an acidic buffer solution.

31. On which of the factors the equilibrium depend : Temperature, nature of reactant and product, initial concentration and pressure of the reactants.

32. Arrange the following in increasing acidic strength HCl, HBr, HF, HI

[**Ans.** HF < HCl < HBr < HI]

33. Arrange the following in increasing Lewis base strength

NH_3, BiH_3, PH_3, AsH_3, SbH_3

[**Ans.** $BiH_3 < SbH_3 < ASH_3 < PH_3 < NH_3$]

34. Arrange following in increasing pH value

0.1M CH_3COOH, 0.1 M NaCl, 0.1MHCl, 0.1MNaOH, 0.1MNH_4OH

[**Ans.** 0.1MHCl < 0.1M CH_3COOH < 0.1M NaCl < 0.1 NH_4OH < 0.1M NaOH]

35. Arrange following in increasing order of degree of hydrolysis.

0.1M NH_4OH, 0.01 M NH_4OH, 10^{-5} M NH_4OH, 10^{-3} M NH_4OH, 10^{-6} M NH_4OH

[**Ans.** 0.1M NH_4OH < 10^{-2} M NH_4OH, 10^{-3}M NH_4OH< 10^{-5}MNH_4OH < 10^{-6} M NH_4OH]

36. Arrange following in increasing order of acidic strength

CH_3COOH, HCOOH, CH_3CH_2COOH, C_6H_5COOH, CH_2COOH

[**Ans.** $CH_3COOH < C_6H_5COOH < HCOOH < CH_2FCOOH$]

37. Arrange the basic strength of following

F^-, Br^-, Cl^-, I^-

[**Ans.** $I^- < Br^- < Cl^- < F^-$]

38. Arrange the following in increasing base strength

CH_3^-, NH_2^-, OH^-, F^-

[**Ans.** $F^- < OH^- < NH_2 < CH_3^-$]

2-MARKS QUESTIONS

1. Calculate pH of 0.001M CH_3COOH having 3% dissociation.

[**Ans.** 4.5229]

2. The equilibrium constant for

$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is K, then calculate equilibrium constant for $\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons NO(g)$. [**Ans.** $\sqrt{K}$]

3. For the reversible reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500°C, the value of Kp is 1.44×10^{-5} atn^{-2}. Find the K_c value.

[**Ans.** $1.44 \times 10^{-5} / (0.082 \times 773)^{-2}$]

4. The equilibrium constant at 298 K for the reaction $A + B \rightleftharpoons C + D$ is 100. If the initial concentration of all the four species were 1M each, then equilibrium concentration of D will be [**Ans.** 1.818]

5. For the reaction

$NH_4COO\ NH_2(s) \rightleftharpoons 2NH_3(g) + CO_2(g)$

If equilibrium pressure is 3 atm. Find the value of Kp [**Ans.** 4]

6. A buffer solution with pH 9 is to be prepared by mixing NH_4Cl that should be added to one litre of 1.0m NH_4OH kb 1.8×10^{-5}

[**Ans.** $NH_4Cl = 1.8$ M]

7. Calculate the solubility of silver chloride in water at room temperature if the K_{sp} of AgCl is 1.6×10^{-10}. [**Ans.** 1.26×10^{-5}M]

8. Calculate the molar solubility of $Ni(OH)_2$ in 0.10m NaOH. The ionic product of $Ni(OH)_2$ is 2.0×10^{-15}. [**Ans.** 2.0×10^{-13}M]

9. Calculate the pH of 10^{-8} M HCl solution. [**Ans.** 6.96]

10. How many grams of NaOH must be dissolved in IL of the solution to give it a pH value of 12? [**Ans.** 0.4g]

3-MARKS QUESTIONS

1. The equilibrium constant for the reaction $H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$ at 1024 K is 1.6×10^5. Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024K. [**Ans.** 10 bar]

2. For the reaction $2BrCl(g) \rightleftharpoons Br_2(g) + Cl_2(g)$ K_c is 32 at 500 K. If initially pure BrCl is present at a concentration of 3.30×10^{-3} M, what is its molar concentration in the mixture at equilibrium? [**Ans.** 3.0×10^{-4} M]

3. When 0.02M HCl solution is added to 0.01M $Pb(NO_3)_2$, will a precipitate of $PbCl_2$ be formed or not. $K_{sp} = 1.6 \times 10^{-5}$ [**Ans:** No]

4. Find the value of Kc for the reaction

   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_2(g)$ $K_p = 3.4$ bar^{-1} at 1000°C

   [**Ans.** 359.24]

5. Ammonium hydrogen sulphide dissociate according to equation

   $NH_4HS(s) \rightleftharpoons NH_3(g) + H_2S(g)$

   If the observed pressure at equilibrium is 1.12 atm at 380K. What is Kp for the reaction [**Ans.** 0.3136]

6. How much of 0.3M NH_4OH should be mixed with 30 mL of 0.2m solution of NH_4Cl to give a buffer solution of pH 10. pk_b for NH_4OH is 4.75.

   [**Ans.** V = 112.5 mL]

7. Predict whether a precipitate will be formed or not on mixing 20 mL of 0.001 N NaCl solution with 80 mL of 0.01 M $AgNO_3$ solution. K_{sp} for AgCl is 1.5×10^{-10}. [**Ans.** Yes, ppt will formed.]

8. The values of Ksp of two sparingly soluble salts $Ni(OH)_2$ and AgCN are 2.0×10^{-15} and 6.0×10^{-17} respectively. Which salt is more soluble. Explain

   [**Ans.** $S_{Ni(OH)_2} = 7.93 \times 10^{-6}M : S_{(Ag\,CN)} = 7.8 \times 10^{-9}M$. $Ni(OH)_2$ is more soluble]

9. The ionization constant of propanoic acid is 1.32×10^{-5}. Calculate the degree of ionization if its solution is 0.05 M. What will be its degree of ionization if the solution is 0.01 M in HCl solution.

   [**Ans.** 1.62×10^{-2}, 1.32×10^{-3}]

10. Calculate the pH of a solution obtained by mixing 50ml of 0.2M HCl with 49.9 mL of 0.2m NaOH solution. [**Ans.** 3.699]

HOTS QUESTIONS

1. The molar solubility of Cd(OH)$_2$ is 1.84 × 10^{-5}M. Calculate the expected solubility of Cd(OH)$_2$ in a buffer solution of pH = 12.

Ans. $Cd(OH)_2 \rightleftharpoons Cd^{2+}_{(aq.)} + 2OH^{-}_{(aq.)}$

$S \qquad 10^{-2}$

$2.49 \times 10^{-14} = S(10^{-2})^2 \qquad \therefore S = 2.49 \times 10^{-10M}$

2. An aqueous solution contains an unknown concentration of Ba^{2+}. When 50 ml of a 1M solution of Na$_2$SO$_4$ is added. BaSO$_4$ just begins to precipitate. The final volume is 500ml. The solubility product of BaSO$_4$ is 1×10^{-10}. Find the original concentration.

Ans. $K_{sp} = \left[Ba^{2+}\right]\left[SO_4^{2-}\right] = \left[Ba^{2+}\right]\left[\frac{50\times1}{500}\right] = 10^{-9}\times500$

$Ba^{2+} = 10^{-9}$ M

$10^{-9} \times 500 = 450 \times M \qquad \therefore M = 1.11 \times 10^{-9}M$

3. An aqueous solution contains 0.10 M H$_2$S and 0.20 M HCl. If the equilibrium constants for the formation of HS$^-$ from H$_2$S is 1.0×10^{-7} and that of S^{2-} from HS$^-$ ions is 1.2×10^{-13}, then find the concentration of S^{-2} ions in aqueous solution.

Ans. $H_2S(aq.) \rightleftharpoons 2H^+ + S^{2-}$

$(0.1 - x) \qquad (2x + 0.29) \qquad x$

$K_a = K_{a_1} \times K_{a_2} = 1.2 \times 10^{-20}$

$1.2 \times 10^{-20} = \frac{(0.2)^2\left[S^{2-}\right]}{0.1}$, $[S^{2-}] = 3 \times 10^{-20}$

4. How many litres of water must be added to 1 litre of an aqueous solution of HCl with a pH of 1 to create an aqueous solution with pH of 2?

Ans. $0.1 \times 1 = (1 + v)(0.01) \Rightarrow v = 9L$

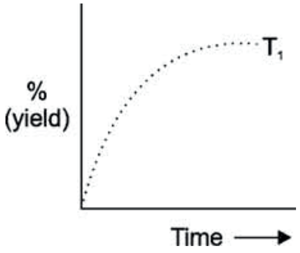
5. A certain buffer solution contains equal concentration of X^- and HX. The K_b for X^- is 10^{-10}. Find the pH of the buffer .

Ans. $k_a.k_b = 10^{-14}$ $\therefore k_a = \frac{10^{-14}}{10^{-10}} = 10^{-4}$

$$pH = pka + \log\frac{[x^-]}{[Hx]}$$

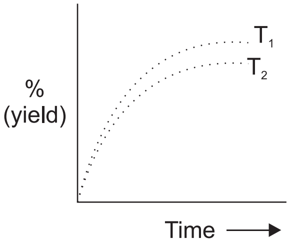
$$\therefore pH = 4 + \log\frac{1}{1} = 4 \qquad \therefore pH = 4$$

6. The % yield of Ammonia as a function of time in the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $\Delta H < O$ at (P, T) is given below:



If this reaction is conducted at $T_2 > T_1$, then plot the % yield of NH_3 as a function of time on same graph

Ans. Initially on increasing temperature the rate of reaction increases, however since the reaction is exothermic therefore % yield of NH_3 get decreased overall after a certain interval of time.



7. Consider the reaction $NH_4COONH_2(s) \rightleftharpoons 2NH_3(g) + CO_2(g)$ at a certain temperature, the equilibrium pressure of the system is 0.318 atm. Find K_p of the decomposition of ammonium carbonate.

Ans. $P_{total} = 3P$ $\therefore P = 0.318/3 = 0.106$

$Kp = 4P^3 = 4.76 \times 10^{-3}$

8. The equilibrium constant for the reaction $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$ is 5. How many moles of CO_2 must be added to 1 litre container already containing 3 moles each of CO and H_2O to make 2M equilibrium concentration of CO?

Ans. $CO + H_2O \rightleftharpoons CO_2 + H_2$

t = 0	3	2	x	0
At equilibrium	2	2	x+1	1

$$\therefore \; S = \frac{x+1}{4} \Rightarrow x = 19$$

9. At constant temperature, the equilibrium constant K_p
$N_2O_4 \rightleftharpoons 2NO_2$ is given by

$$k_p = \frac{4x^2P}{1-x}$$ where, P = Pressure and X = Extent of reaction

How does the value of K_p change on following changes

(a) 'P' increases

(b) X changes

(c) 'P' decreases

Ans. K_p is equilibrium constant which does not change on changing the P, x.K_p depends on temperature.

10. When two reactants A and B are mixed to give product 'c' and 'p' the reaction quotient 'Q' at the initial stages of the reaction will be?

Ans. In the beginning of the reaction Q = 0. As the reaction proceeds in the forward direction Q starts increasing.

At chemical equilibrium Q = K

UNIT TEST-I

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. What is the pH 10^{-3} M HCl solution ? [1]

   (a) 1 (b) 11 (c) 3 (d) 14

2. Which one can act as Arrhenius Acid ? [1]

   (a) NH_3 (b) H_2O (c) HCl (d) C_6H_5OH

3. Write the conjugate base of CH_3COOH, H_2O. [1]

4. Write the relation between K_p and K_c. [1]

5. What is the nature of following reaction [1]

   Exothermic or endothermic

   $A + B - 70J \longrightarrow C$

6. The pka of CH_3COOH and pkb of NH_4OH are 4.76 and 4.75 respectively. Calculate the pH of CH_3COONH_4. [2]

7. What is a buffer solution. Calculate the pH of the solution obtained by adding 4mol of CH_3COOH with 3 mol of NaOH in 1 litre container. [2]

   pka, $CH_3COOH = 4.74$ $\log 2 = 0.3010$ $\log 3 = 0.4771$

8. Calculate the molar solubility of $Ni(OH)_2$ in 0.1M KOH solution. The K_{sp} for $Ni(OH)_2$ is 2.0×10^{-15}. [3]

9. $K_p = 0.04$ atm at 899 K for the equilibrium shown below. What is the equilibrium concentration of H_2 when it is placed in a flask at 4.0 atm pressure and allowed to come to equilibrium. [3]

   $C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$

10. The first ionization constant of H_2S is 9.1×10^{-4}. Calculate the concentration of HS^{-1} in its 0.1 M solution. How will this concentration be effected if the solution is 0.1 M HCl also? If the second dissociation constant of H_2S is 1.2×10^{-12}. Calculate the concentration of S^{2-} in both conditions. [5]

UNIT TEST-II

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. What is the pH 10^{-8} MHCl solution ? [1]

   (a) pH>7 (b) pH<7

   (c) pH=7 (d) Cannot be defined

2. What is the conjugate acid for the NH_3? [1]

   (a) NH_2^- (b) NH^{2-}

   (c) N^{3-} (d) NH_4^+

3. Define Lewis acid and base with one example each. [1]

In following questions a statement of assertion followed by a statement of reason is given. Choose the correct answer out of the following choices.

(a) Assertion and Reason both are correct statement and reason is correct explanation for assertion.

(b) Assertion and Reason both are correct statements but reason is not correct explanation for assertion.

(c) Assertion is correct statement but reason is wrong statement.

(d) Assertion is wrong statement but reason is correct statement.

4. **Assertion :** HI is stronger acid than HCl [1]

   Reason : Size of I is more than Cl and hence H-I bond strength is less than HCl.

5. **Assertion :** BF_3 is Lewis acid and NH_3 is Lewis base. [1]

   Reason : NH_3 is short of octet and BF_3 molecule contain lone pair of electron.

6. Arrange the following in increasing acidic strength. Give reason also

   CH_4, NH_3, H_2O, HF [2]

7. K_c for the reaction $SO_2 + 0.5O_2 \rightarrow SO_3$ at 600°C is 61.7. Calculate K_p. [2]

8. 25.4 ml of hydrogen and 20.4 ml of iodine when heated in a closed container, produced 30.8 mL of HI at equilibrium. Calculate the degree of dissociation of HI at same temperature. [3]

9. Define common ion effect. The solubility of CaF_2 in water at T K is $2x10^{-4}$ moles/L. Calculate (i) K_{sp}, and (ii) Solubility in 0.01 M NaF solution. [3]

10. (i) What is a buffer solution? What are its types?

    (ii) Derive the Henderson-Hasselbalch equation for an Acidic buffer with the help of relevant example.

    (ii) 8g of NaOH was dissolved in one litre of a solution containing one mole of CH_3COOH and one mole of CH_3COONa. Find the pH of the resulting solution. (The pK_a of CH_3COOH is 4.74). [5]

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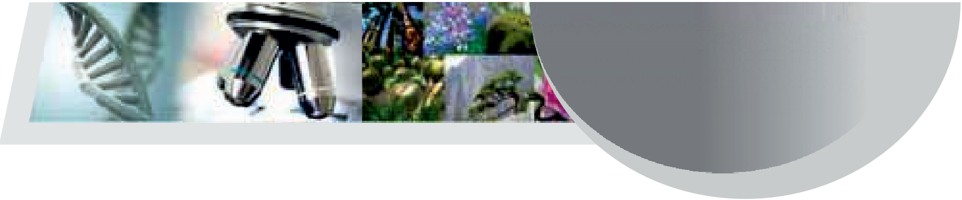
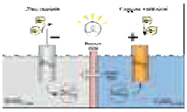




Chapter - 7

Redox Reactions

FAST TRACK : QUICK REVISION

Oxidation and Reduction :

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

- **Reducing Agent :** Reduce other substance and oxidise itself.
- **Oxidising Agent :** Oxidise other substance but reduce itself.
- **Redox Reaction :** Reactions in which oxidation and reduction takes place simultaneously.
- **Oxidation Number :** It is charge that an atom appears to have in a given species when the bonding electron are counted towards more electro-negative atom.
- **Calculation of Oxidation Number :**
  - (a) Oxidation number of all the elements in their elemental form (in standard state) is taken as zero. Oxidation number of element in a molecule Cl_2, F_2, O_2, P_4, O_3, Fe, H_2, N_2, C (graphite) is zero.
  - (b) Common Oxidation number of elements of first group is +1. Common Oxidation number of elements of second group + 2.
  - (c) For ions composed of only one atom, the oxidation number is equal to the charge on the ion.

(d) The oxidation number of oxygen in most compounds is – 2. While in peroxides (*e.g.*, H_2O_2, Na_2O_2), each oxygen atom is assigned an oxidation number of – 1, in super oxides (*e.g.*, KO_2, RbO_2) each oxygen atom is assigned an oxidation number of – (½).

(e) In oxygen difluoride (OF_2) and dioxygen difluoride (O_2F_2), the oxygen is assigned an oxidation number of + 2 and + 1, respectively.

(f) The oxidation number of hydrogen is + 1 but in metal hydride its oxidation no. is – 1.

(g) In all its compounds, fluorine has an oxidation number of – 1.

(h) The algebraic sum of the oxidation number of all the atoms in a compound must be zero.

(i) In polyatomic ion, the algebraic sum of all the oxidation numbers of atoms of the ion must equal the charge on the ion.

- **Types of Redox Reactions:**

(i) Combination Reaction : 0 0 + 2 – 3

$$3\,Mg\,(s) + N_2\,(g) \xrightarrow{\Delta} Mg_3N_2\,(s)$$

(ii) Decomposition Reaction : +1 +5 – 2 + 1 – 1 0

$$2KC1O_3(s) \xrightarrow{\Delta} 2\,KC1(s) + 3O_2\,(g)$$

(iii) Metal Displacement : + 2 + 6 – 2 0 + 2 + 6 – 2 0

$$CuSO_4\,(aq) + Zn(s) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$$

(iv) Non-metal displacement : 0 + 1 – 2 + 2 – 2 + 1 0

$$Ca\,(s) + 2\,H_2O\,(1) \rightarrow Ca\,(OH)_2 + H_2\,(g)$$

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

0 –1 + 1

$$C1_2\,(g) + 2\,OH^-\,(aq) \rightarrow Cl^-\,(aq) + ClO^-\,(aq) + H_2O\,(1)$$

- **Stock Notation :** Representing oxidation number of metal in Roman numerals within parenthesis after the symbol or name of metal in the molecular formula or name of a compound. For *e.g.*, Stock Notation of Ferric oxide is $Fe_2(III)O_3$ or Iron (III) oxide.

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

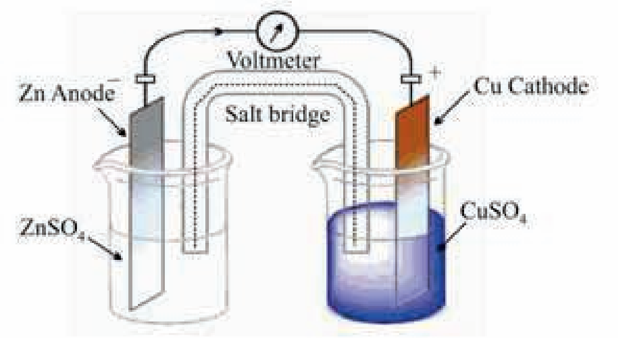
In reality no element can have a fractional oxidation state.

- **Balancing of Redox Reactions :**

  (A) Oxidation number method

  (B) Half reaction method

- **Electrode Potential (E) :** Potential difference between electrode and electrolytic solution due to charge separation.
- **Standard Electrode Potential (E^θ) :** Electrode Potential measured at 298 K and 1M concentration of metal ions (or 1 bar pressure of gas).
- **Electrochemical Cell :** A device in which chemical energy of a spontaneous redox reaction is converted into electrical energy.



Cell diagram: Zn | Zn^{2+} || Cu^{2+} | Cu

LHS oxidation, $Zn \rightarrow Zn^{2+} + 2e^-$

RHS reduction $Cu^{2+} + 2e^- \rightarrow Cu$

Overall reaction $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

- Representation of an Electrochemical cell :

$$\xrightarrow{\text{Flow of electrons}}$$

$$\xleftarrow{\text{Flow of current}}$$

$$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

	Left Electrode	Salt Bridge	**R**ight Electrode
LOAN	**O**xidation		**R**eduction
	Anode		**C**athode
	Negative		**P**ositive

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

MIND MAP : REDOX REACTIONS

REDOX REACTION

REDUCTION

- Addition of Hydrogen
  $Na + H_2 \longrightarrow NaH$
- Addition of electropositive element
  $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$
- Removal of Oxygen
  $2HgO \longrightarrow 2Hg(l) + O_2(g)$
- Removal of electronegative element
  $2FeCl_3 + H_2 \longrightarrow 2FeCl_2 + 2HCl$
- Gain of electron or decrease in oxidation number
  $2Na + Cl_2 \longrightarrow 2\overset{+}{Na}Cl^-$
  Cl get reduced

Oxidant or Oxidising Agents
Substance which get reduced and help to oxidise other substance.

OXIDATION

- Addition of oxygen
  $Mg + O_2 \longrightarrow MgO$
- Addition of electronegative element
  $Mg + F_2 \longrightarrow MgF_2$
- Removal of Hydrogen
  $2H_2S + O_2 \longrightarrow 2S + 2H_2O$
- Removal of electronegative element
  $2K_4[Fe(CN)_6] + H_2O_2 \longrightarrow 2K_3[Fe(CN)_6] + 2KOH$
- Loss of electron or increase in oxidation number
  $2Na + Cl_2 \longrightarrow Na^+Cl^-$
  Na — oxidise

Reduction / Reducing Agent
Substance which get oxidised itself but help to reduce others substance is called Reducing Agent / Reductant.

Types of Redox Reaction

- Combination Reactioin
  $\overset{0}{C} + \overset{0}{O_2} \longrightarrow C^{+4}O_2^{-2}$
- Decomposition Reaction
  $2K\overset{+5}{Cl}O_3^{-2} \longrightarrow 2KCl^{-1} + 3\overset{0}{O_2}$
- Displacement Reaction
  a) Metal Displacement
  $Cr_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Cr$
  b) Non-Metal Displacement
  $2Na + 2H_2O \longrightarrow 2NaOH + H_2$
- Disproportionation Reaction
  $H_2O_2 \longrightarrow 2H_2O + O_2$

ELECTROCHEMICAL

$Zn\,|Zn^{2+}|\,|Cu^{2+}|Cu$

$E^\circ_{Cell} = E - E^\circ_L = E^\circ_{Cathode} - E^\circ_{Anode}$

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

$$E_{Cell} = E^\circ_{Cell} - \frac{0.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

E_{Cell} = 0 At equilibrium

E_{Cell} < 0 Non spontaneous reaction

E_{Cell} > 0 Spontaneous reaction

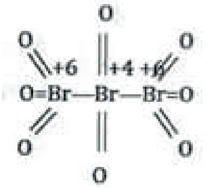
CASE BASED STUDY - QUESTIONS

1. Read the given passage and answer the questions.

Redox reactions are reactions in which oxidation and reduction takes place simultaneously. Oxidation number are assigned in accordance with the set of rules. Oxidation number and ion electron methods both are used in balancing ionic equations. Redox reactions are classified as combination, decomposition, displacement and disproportionation reactions. The concept of redox couple and electrode processes is basis of electrolysis and electrochemical cells.

(a) What are oxidation number of each individual Br in Br_3O_8

Ans. +6, +4, +6

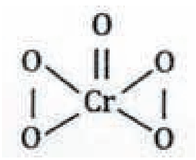


(b) If electrolysis of $CuSO_4$, solution is carried out using Cu electrodes, what will be reaction taking place at anode.

Ans. $Cu \rightarrow Cu^{2+} + 2e^-$

(c) What is oxidation number of Cr in CrO_5?

Ans. +6



∵ It has peroxide linkage.

(d) Give one example of disproportionation reaction.

Ans. $2Cu^+ \rightarrow Cu^{2+} + Cu$

(e) $MnO_4^{2-} + H^+ \rightarrow MnO_4^- + MnO_2 + H_2O$ [Balance this reaction]

Ans. $MnO_4^{2-} \rightarrow MnO_4^- + e^-$ ....(i)×2

$2e^- + 4H^+ + MnO_4^{2-} \rightarrow MnO_2 + 2H_2O$ ....(ii)

$3MnO_4^{2-} + 4H^+ \rightarrow 2MnO_4^- + MnO_2 + 2H_2O$

2. Redox Reactions : Passage Based Question (Assertion and Reason)

Passage : Redox reactions are those reactions in which, there is a simultaneous oxidation and reduction taking place. There is an addition of oxygen and removal of hydrogen taking place in oxidation reactions. In Reduction, hydrogen gets added and oxygen gets removed. Redox reactions are also used to determine the strength of reductant/oxidant. In oxidation, there is a decrease in electron density while in reduction, there is an increase in electron density around the atom.

(Q1-Q4) There are assertion and reason which have been put forward. Read the given statement and choose correct alternative from the following :

(Note : A-Assertion & R-Reason)

(a) Both A and R are true and R is the correct explanation of A.

(b) Both A and R are true but R is not the correct explanation of A.

(c) A is true but R is false.

(d) A is false but R is true.

Q1. **A :** Oxidation-Reduction (Redox) couple is the combination of oxidized and reduced form of a substance that is involved in Oxidation-Reduction half cell.

R: As in representation E° Fe^{3+}/Fe^{2+} and E° Cu^{2+}/Cu^{+} are two Redox couples.

Q2. **A :** In Formaldehyde (HCHO) oxidation no. of carbon is 0.

R: Formaldehyde is a covalent compound.

Q3. **A :** Oxidation state of hydrogen in water is +1 and CaH_2 is -1.

R: CaH_2 is metal hydride and for hydrogen, it assigned the oxidation state of -1.

Q4. **A :** Redox reactions are also called neutralization reactions.

R: As the number of electron gained/lost in the reaction are balanced.

Or

A : Substances which get reduced act as reducing agent.

R: Oxidizing agent itself gets reduced.

Ans. : Q1. (c), Q2. (b), Q3. (a), Q4. (d) or (d)

MULTIPLE CHOICE QUESTIONS (MCQ)

1. The oxidation state of Fe in Fe_3O_4 is

(a) +2
(b) + 3
(c) $\frac{8}{3}$
(d) + 2, + 3

2. The oxidation state of 'S' in $KAl(SO_4)_2 . 12H_2O$ is

(a) –2
(b) – 1
(c) 2
(d) + 6

3. Oxidation state carbon in C_3O_2 is

(a) $\frac{4}{3}$
(b) 0
(c) 2
(d) 0, 2

4. The reaction $S_8 + 12OH^- \longrightarrow 4S^{2-} + 2S_2O_3^{2-} + 6H_2O$ is

(a) Combination reaction
(b) Decomposition reaction
(c) Non-metal displacement
(d) Disproportionation reaction

5. E^0 for H^+/H_2 is

(a) 0
(b) +1V
(c) –1.0V
(d) –2.0V

6. Which one act as strong oxidising agent

$K^+/K = -2.9\ 3V, Ag^+/Ag = 0.80, Hg^{2+}/Hg = 0.79V$

(a) K^+
(b) K
(c) Hg^{2+}
(d) Ag^+

7. The coefficient of HCl in balance reaction is

$Pb_3O_4 + HCl \longrightarrow PbCl_2 + Cl_2 + H_2O$

(a) 1
(b) 8
(c) 3
(d) 4

8. Sum of oxidation numbers of all Bromine atoms in Br_3O_8 is

(a) 6
(b) 4
(c) 16
(d) 20

9. In the reaction $6ClO_2^- \longrightarrow 4ClO_3^- + 2Cl^-$, Cl^- ion is

(a) Oxidised Reduced
(b) Reduced
(c) Odixised and
(d) Neither Oxidised nor reduced

10. 'I' can not act as reducing agent in following state

(a) –1 (b) +1

(c) +7 (d) +5

Ans: 1. (d) 2. (d) 3. (d) 4. (d) 5. (a) 6. (d) 7. (b) 8. (c) 9. (c) 10. (d)

FILL IN THE BLANKS

(i) Oxidation is __________ of electrons.

(ii) S.H.E. stands for __________.

(iii) Oxidation state of Oxygen in O_2F_2 is __________.

(iv) Disproportionation is a type of __________ reaction.

(v) Oxidant is one which __________ electron..

(vi) $Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^-$ is a __________ type of reaction.

(vii) Oxidation state of F is always either ______ or _____ .

(viii) Oxidation state of Oxygen in O_3 is __________ .

(ix) Reducing agent are also called __________.

(x) Hydrogen economy is use of Hydrogen as _________.

Ans: (i) loss, (ii) standard hydrogen electrode, (iii) +1, (iv) redox, (v) gain, (vi) disproportionation, (vii) 0, –1, (viii) zero, (ix) reductant, (x) fuel

TRUE AND FALSE TYPE QUESTIONS

(i) In Redox reaction first oxidation take place.

(ii) Oxidising agents are also called reductant.

(iii) Fluorine cannot have +1 oxidation state.

(iv) O_2^+ has oxidation state of oxygen as +½.

(v) If for the reaction $Ca^{2+} + 2e^- \longrightarrow Ca(s)$; $E^\ominus = -2.87$

Then for the reaction $2Ca^{2+} + 4e^- \longrightarrow 2Ca(s)$; $E^\ominus = 2(-2.87)V$

(vi) Salt bridge is used for enhancing E° values of individual half reaction.

(vii) Anode is –ve charged in Galvanic cell.

(viii) KCl can be use in salt bridge.

(ix) Current flows in galvanic cell from Anode to cathode.

(x) MnO_4^- is colourless in basic medium.

Ans: (i) False (ii) False (iii) True (iv) True (v) False
(vi) False (vii) True (viii) True (ix) False (x) False

MATCH THE COLUMNS

1.

Column-I	Column-II
(a) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$	(p) Disproportionation
(b) $Cr_2O_3 + 2A1 \longrightarrow Al_2O_3 + 2Cr$	(q) Non-metal displacement
(c) $Fe + 2HCl \longrightarrow FeCl_2 + H_2$	(r) Metal displacement
(d) $P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$	(s) Metal displacement

Ans. (a) – (s) (b) – (r) (c) – (q) (d) – (p)

2.

Column-I	Column-II (Oxidation sate of N)
(a) NO	(p) + 5
(b) NO_2	(q) + 3
(c) NO_2^-	(r) + 4
(d) NO_3^-	(s) + 2

Ans. (a) – (s) (b) – (r) (c) – (q) (d) – (p)

3.

Column-I	Column-II
(A) Increase in oxidation number	(a) Loss of electron
(B) Reducing agent	(b) Oxidation
(C) $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + H_2O$	(c) Natural redox reaction
(D) Photosynthesis	(d) Redox reaction

Ans. (A) – (b) (B) – (a) (C) – (d) (D) – (c)

4.

Column-I	Column-II
(A) Decrease in oxidation number	(a) Disproportionation
(B) Oxidizing agent	(b) Fractional oxidation number
(C) $2Cu^+ \rightarrow Cu^{2+} + Cu$	(c) Reduction
(D) Mn_3O_4	(d) Gain of electron

Ans. (A) – (c) (B) – (d) (C) – (a) (D) – (b)

5.

Column-I	Column-II
(A) H_2O_2	(a) –1
(B) $MnSO_4$	(b) +3
(C) $AlCl_3$	(c) +5
(D) P_2O_5	(d) +6

Ans. (A) – (a) (B) – (d) (C) – (b) (D) – (c)

ASSERTION AND REASON TYPE QUESTIONS

Each question contain statement-1 (Assertion) and statement-2 (Reason)

Examine the statements carefully and work the correct answer according to the instruction given below :

(a) If both the statements are True and Statement-2 is the correct explanation of the statement-1

(b) If both the statements are True and statement-2 is not the correct explanation of statement-1

(c) If statement-1 is true and statement-2 is False.

(d) If statement-1 is false and statement-2 is True.

1. Statement - 1 : In HF, the oxidation state of 'F' is –1

   Statement - 2 : 'F' being most electronegative, will have -1 oxidation in its compound.

2. Statement - 1 : Oxygen has zero oxidation state in O_2.

   Statement - 2 : Element in their elemental form have zero oxidation state.

3. Statement - 1 : Oxidation state of Oxygen in H_2O_2 is –1.

   Statement - 2 : H_2O_2 has peroxide linkage.

4. Statement - 1 : For the reaction $Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$; E_{cell} is +ve.

   Statement - 2 : For standard Hydrogen electrode $E° = 0$

5. Statement - 1 : $2H_2O_2 \longrightarrow 2H_2O + O_2$ is Decomposition reaction (Redox).

   Statement - 2 : Oxygen has -2 oxidation state in H_2O.

6. Statement - 1 : $C + O_2 \longrightarrow CO_2$; carbon get oxidised.

   Statement - 2 : Gain of Hydrogen is reduction.

7. Statement - 1 : $CaCO_3 \longrightarrow CaO + CO_2$ is not redox reaction.

   Statement - 2 : C, Ca, O do not change their oxidation number in the reaction.

8. Statement - 1 : Oxidation also occurs when decrease in electron density is observed.

   Statement - 2 : Oxidation is gain of electro-positive element.

9. Statement - 1 : $Cr_2O_7^{2-}$ is a self indicator.

   Statement - 2 : MnO_4^- acts as a self indicator.

10. Statement - 1 : Equivalence point comes first before end point.
Statement - 2 : Equivalence point cannot be obtained even by graphical method.

Ans: 1. (a) 2. (a) 3. (a) 4. (b) 5. (d) 6. (b) 7. (a) 8. (c) 9. (d) 10. (d)

ONE WORD ANSWER TYPE QUESTIONS

1. What is the oxidation number of S in S_8.
2. What is the oxidation state of Oxygen in H_2O_2.
3. Name the substance used in salt-bridge.
4. Name an indicator which can act as self-indicator.
5. When a substance gains electron, it is called :
6. Name the ion which is used for balancing the hydrogen atom in acidic medium.
7. In the reaction $3Mg + N_2 \longrightarrow Mg_3N_2$, Nitrogen is oxidised or reduced.

Ans: 1. zero 2. 1 3. NH_4Cl or KCl 4. $KMnO_4$
5. Reduction 6. H^+ 7. Reduced

1-MARK QUESTIONS

1. Define oxidation and reduction according to electronic concept.

2. Define oxidation and reduction according to oxidation number.

3. A freshly cut apple is almost white but it turns reddish brown after sometime. Give reason.

4. Define oxidation number.

5. Write oxidation number of Mn in $KMnO_4$.

6. Write oxidation number of Cr in $Cr_2O_7^{2-}$.

7. Write Stock notation of MnO_2 and $AuCl_3$.

8. Define redox reaction with example.

9. Define disproportionation reaction. Give one example.

10. Define the term redox titration.

11. Name the indicator used in redox titration involving $K_2Cr_2O_7$ as an oxidizing agent.

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

13. The standard reduction potentials of three metals cations X, Y and Z are + 0.52, – 3.03 and – 1.18 V respectively. Arrange X, Y and Z in order of increasing reducing power. [**Ans.** $X < Z < Y$]

14. An electrochemical cell consists of two electrodes *i.e.*, Anode and Cathode. What is the direction of flow of electrons in this cell ?

15. Why anode is negatively charged in an electrochemical cell?

16. Out of Zn and Cu vessel one will be more suitable to store 1 M HCl?

[**Ans.** Cu]

Given $E^{\theta}_{Zn^{2+}/Zn} = -0.76$ V, $E^{\theta}_{Cu^{2+}/Cu} = +0.34$ V.

15. Is it safe to stir 1 M $AgNO_3$ solution with copper spoon ?

Given $E^{\theta}_{Ag^{+}/Ag} = +0.80$ V, $E^{\theta}_{Cu^{2+}/Cu} = +0.34$ V. [**Ans.** No]

2-MARKS QUESTIONS

1. Identify oxidant and reductant in the reaction :

$$I_2\,(aq) + 2S_2O_3^{2-}\,(aq) \longrightarrow 2\,I^-\,(aq) + S_4O_6^{2-}(aq).$$

2. Calculate oxidation number of Fe in Fe_3O_4 and write a suitable justification of your answer.

3. Oxidation-reduction reactions are complementary. Explain.

4. Write formula for the following compounds :

(i) Mercury (II) chloride

(ii) Nickel (II) sulphate

(iii) Iron (III) sulphate

(iv) Chromium (III) oxide

5. Justify that the reaction : $H_2O(s) + F_2 \longrightarrow HF + HOF$ is a redox reaction.

6. A decomposition reaction may or may not be a redox reaction. Write two decomposition reactions in support of the statement.

7. Split the reaction $2\ K\ (s) + C1_2\ (g) \longrightarrow 2\ KC1\ (s)$ into oxidation and reduction half reactions.

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

   (i) $Ca\underline{O}_2$ (ii) $H_2\underline{S}_2O_7$ (iii) $K_2\underline{Mn}O_4$ (iv) $K\underline{I}_3$

9. Write the functions of salt bridge in an electrochemical cell.

10. Define the term redox couple. Write the practical application of redox couple.

11. The standard reduction potentials of two metals A and B are – 0.76 V and + 0.34 V respectively. An electrochemical cell is formed using electrodes of these metals.

    (i) Identify the cathode and anode.

    (ii) Write the direction of flow of electron.

3-MARKS QUESTIONS

1. Calculate oxidation number of :

   (i) Cr in $Cr_2O_4^{2-}$

   (ii) O in KO_2

   (iii) Na in Na_2O_2.

2. Account for the following :

   (i) HNO_3 acts as oxidizing agent while HNO_2 can act both as reducing and oxidizing agent.

   (ii) AgF_2 is unstable compound and act as a strong oxidizing agent.

   (iii) Ozone acts as an oxidising agent.

3. Permanent ion (MnO_4^-) reacts with sulfur dioxide gas in acidic medium to produce Mn^{2+} ion and hydrogen sulphate ion. Write ionic equation and balance by ion electron method.

4. Balance the following equation by oxidation number method :

   $P_4(s) + OH^-(aq) \longrightarrow PH_3 + H_2PO_2^-(aq)$ [*Basic Medium*]

5. Balance the following equation by ion electron method :

   $C1_2O_7(g) + H_2O_2(1) \longrightarrow C1O_2^-(aq) + O_2(g)$ [*Basic medium*]

6. Depict the galvanic cell in which the reaction

   $Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Further show :

   (i) Which electrode is negatively charged ?

   (ii) The carriers of the current in the cell

   (iii) Individual reaction at each electrode.

7. Explain with suitable reasons :

   (i) Reaction $FeSO_4(aq) + Cu(s) \longrightarrow CuSO_4(aq) + Fe$ does not occur.

   (ii) Zinc can displace copper from aqueous $CuSO_4$ solution but Ag cannot.

   (iii) Solution of $AgNO_3$ turns blue when copper rod is immersed in it.

5-MARKS QUESTIONS

1. (i) MnO_4^{2-} undergoes disproportionation reaction in acidic medium but MnO_4^- does not. Give reason.

   (ii) Give one example each of the following redox reactions:

   (a) Combination reaction

   (b) Decomposition reaction

   (c) Metal displacement reaction

2. Consider the cell reaction of an electrochemical cell : $Ni(s) + 2Ag^+(aq) \rightarrow Ni^{2+}(aq) + 2Ag(s)$ and answer the following questions :

   (i) Write anode and cathode half reactions.

   (ii) Mention the direction of flow of electrons.

(iii) How is the electrical neutrality maintained in the solutions of the two half cells ?

(iv) Write the formula for calculating standard emf of this cell.

(v) How does the emf change when the concentration of silver ions is decreased ?

3. Justify the reason that following reactions are redox reactions.

(a) $CuO\ (s) + H_2\ (g) \longrightarrow Cu\ (s) + H_2O\ (g)$

(b) $Fe_2O_3 + 3CO\ (g) \longrightarrow 2\ Fe\ (g) + 3CO_2\ (g)$

(c) $NH_3\ (g)\ 5O_2\ (g) \longrightarrow 4NO\ (g) + 5H_2O\ (g)$

(d) $BCl_3\ (g) + 3\ LiAlH_4 \longrightarrow B_2H_6 + LiCl + AlCl_3$

(e) $2K + F_2 \longrightarrow 2KF$

[**Hints:–** CuO is oxidizing agent, H_2 is acting as reducing agent because Cu (II) is changing to Cu (0) by gain of e^-,H_2 is getting oxidised to H_2O (g), its oxidations sate is changing from 0 to +1, by loss of electrons.

(ii) It is redox reaction: Fe_2O_3 is getting reduced to fe. CO is getting oxidised to CO_2.]

4. Using standard electrode : Predict if the reaction between as the following is feasible.

(i) Fe^{3+} (aq) and I^- (aq)

(ii) Ag^+ and Cu

(iii) Fe^{3+} and Br^- (aq)

(iv) Ag and Fe^{3+} (aq)

(iv) Br_2 (aq) and Fe^{2+} (aq)

Hint:– $E^\theta_{I_2/I^-} = 0.541$ V, $E^\theta_{Cu^{2+}/Cu'} = 0.34V$, $E^\theta_{Br_2/Br^-} = 1.09V$, $E^\theta_{Ag^+/Ag} = 0.80V$, $E^\theta_{Fe^{3+}/Fe^{2+}} = 0.77V$.

5. Draw the diagram for the galvanic cell which would have overall chemical reaction as

$$Zn + 2Ag^+ \longrightarrow Zn^{2+} + 2Ag.$$

Answer the following :

(i) Write the reactions occurring at each electrode.

(ii) In which directions do the electrons flow in the external circuit?

(iii)Name the salt to be taken in salt bridge.

(iv)Label the anode and cathode.

(v) How does the EMF change when the concentration of solvers ions is decreased?

HOTS QUESTIONS

1. 6×10^{-3} mole $K_2Cr_2O_7$ reacts completely with 9×10^{-3} mole X^{n+} to give XO_3^- and Cr^{3+}. Find the value of X.

Ans. $K_2Cr_2O_7 + X^{n+} \longrightarrow \overset{+5}{X}O_3^- + Cr^{3+}$

$6\times10^{-3} \times 6 = (5-n) \times 9 \times 10^{-3} \longrightarrow n = 1$

2. For the redox reaction

$K_2Cr_2O_7 + X\, H_2SO_4 + Y\, SO_2 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + ZH_2O$

What is the sum of $x + y + z$

Ans. $K_2Cr_2O_7 + H_2SO_4 + 3SO_2 \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O$

$\therefore x = 1 \quad y = 3 \quad z = 1 \quad \therefore x + y + z = 5$

3. An aqueous solution containing 1M each of Au^{+3}, Cu^{+2}, Ag^+, Li^+ is being electrolysed using inert electrodes the value of standard potentials are

$E^{\theta}_{Ag^+/Ag} = 0.80$ V, $E^{\theta}_{Cu^{2}+Cu} = 0.34$ V, $E^{\theta}_{Au^{3+}/Au} = 1.50$ V, $E^{\theta}_{Li^+/Li} = -3.03$V

With increasing voltage, find the sequence of deposition of metals on the cathode.

Ans. Only Au^{3+}, Ag^+ and Cu^{2+} will deposit at cathode.

Li^+ will not deposit at cathode be cause SRP of water is -0.8274V

So after Cu^{2+} ; H_2 will evolve at cathode.

4. E^{θ} for $Cl_2 (g) + 2I^- \longrightarrow 2Cl^- (aq.)$ is 1.36 V, then calculate.

E^{θ} for $4C1^- (aq.) \longrightarrow 2Cl_2(g) + 4e^-$

Ans. $E^{\theta}_{Cl^-/Cl_2} = -1.36$ E^{θ} is independent of amount of substance

5. Why salt bridge is made up of saturated solution of KNO_3 in agar–agar.

Ans. Velocities of both K^+ and NO_3^- are nearly the same.

UNIT TEST-I

Time Allowed: 1 hr **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. Identify the oxidised and Reduced species in the following reaction [1]

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

(a) H_2S (b) Cl_2 (c) Both H_2, Cl_2 (d) None of these

2. What is the oxidation state of Br in BrO_3^-? [1]

(a) +1 (b) +3 (c) +4 (d) +5

3. Classify the type of reaction in Redox Reaction form : [1]

$$3H_2O + P_4 + 3OH^- \longrightarrow PH_3 + 3H_2PO_2^-$$

4. What is a redox couple? Give one example.

5. Identify oxidant in reaction given below : [1]

$$CuO(s) + H_2\,(g) \longrightarrow Cu(s) + H_2O(g)$$

6. Assign oxidation number to the underlined elements [2]

(a) $NaH_2\,\underline{P}O_4$ (b) $H_4\,\underline{P_2}O_7$ (c) $K_2\,\underline{Mn}O_4$ (d) $H_2\,\underline{S_2}O_7$

7. Predict product of electrolysis in following case [2]

- An aqueous solution of $CuCl_2$ with platinum electrodes.

8. Consider the reaction $Zn(s) + 2Ag^+\,(aq.) \longrightarrow Zn^{2+}\,(aq.) + 2Ag(s)$

Answer following : [3]

(i) Which electrode is negatively charged ?

(ii) What are carrier of current in the cell ?

(iii) Individual reaction at each electrode.

9. E^θ values are given : $K^+/K = -2.93V$, $Ag^+/Ag = 0.80V$ [3]

$Hg^{2+}/Hg = 0.79V$ $Mg^{2+}/Mg = -2.37V$, $Cr^{3+}/Cr = -0.74V$

(i) Which one is strong reducing agent ?

(ii) Which one is strong oxidising agent ?

(iii) Which redox couple is a stronger reducing agent than H^+/H_2 ?

10. Balance the reaction (ion-electron or oxidation number) [5]

$$P_4(s) + OH^-(aq.) \longrightarrow PH_3\,(g) + H_2PO_2^-\,(aq.) \text{ [Basic medium]}$$

UNIT TEST-II

Time Allowed: 1 Hr. (REDOX REACTIONS) Maximum Marks : 20

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. The average oxidation No. of Iodine is I_3^- ion is. [1]

2. What is oxidation state of Cr in $K_2Cr_2O_7$? [1]

3. Write the name of cell in which chemical energy is converted into Electrical energy. [1]

4. Why is anode negatively charged in an electrochemical cell? [1]

5. Identify the oxidised and Reduced species in the following reaction

   $H_2S + Cl_2 \longrightarrow 2HCl + S$ [1]

6. A decomposition reaction may or may not be a redox Reaction. Write two decomposition reactions in support of the statement. [2]

7. Write the functions of salt bridge in a electrochemical cell. [2]

8. Account for the following : [3]

   (i) HNO_3 acts as oxidizing agent while HNO_2 can act both as Reducing and oxidizing agent.

   (ii) AgF_2 is unstable compound and act as a strong oxidizing agent.

   (iii) Ozone acts as an oxidizing agent.

9. Explain with suitable reasons : [3]

   (i) Reaction $FeSO_4(aq) + Cu(s) \rightarrow CuSO_4(aq) + Fe$ does not occur.

   (ii) Zinc can displace copper from aqueous $CuSO_4$ solution but Ag cannot.

   (iii) Solution of $AgNO_3$ turns blue when copper rod is immersed in it.

10. (i) Give one example each of the following redox reactions : [3]

    (a) Combination Reaction

(b) Decomposition Reaction

(c) Metal displacement Reaction

(ii) Remaining two are Assertion and Reason. Read the statement carefully and choose the correct alternative : [2]

(a) Both Assertion and Reason are true and Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are true but Reason is not the correct explanation of Assertion.

(c) Assertion is true but Reason is false.

(d) Both Assertion and Reason are false.

(I) **Assertion :** Redox couple is the combination of oxidised and Reduced from of a substance involved in an oxidation or reduction half cell.

Reason : In the Representation E^θ Fe^{3+}/Fe^{2+} and E^θ Cu^{2+}/Cu, Fe^{3+}/Fe^{2+} and Cu^{2+}/Cu are Redox couples.

(II) **Assertion :** Oxidation no. of C in HCHO is zero.

Reason : Formaldehyde is covalent compound.

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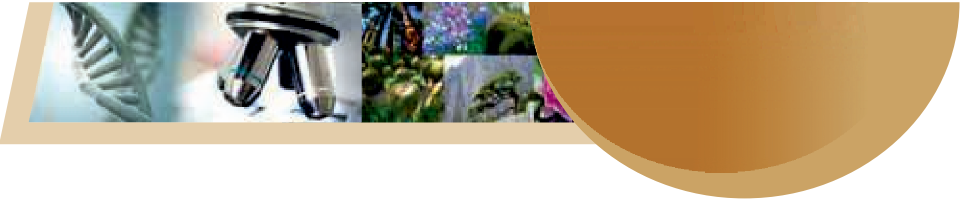
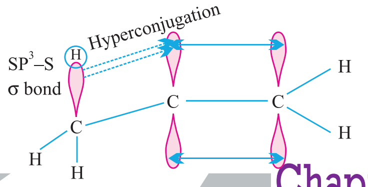




Chapter - 8

Organic Chemistry : Some Basic Principles and Techniques

FAST TRACK : QUICK REVISION

ORGANIC CHEMISTRY

It deals with the study of hydrocarbons (compounds of carbon and hydrogen elements) and their derivatives.

Some organic compounds may also contain nitrogen, oxygen, sulphur, phosphorus, halogens, etc.

Berzelius, proposed that a 'vital force' was responsible for the formation of organic compounds.

This was rejected by F. Wohler who synthesised first organic compound urea from an inorganic compound.

$$\underset{\text{Ammonium cyanate}}{NH_4CNO} \xrightarrow{\Delta} \underset{\text{Urea}}{NH_2CONH_2}$$

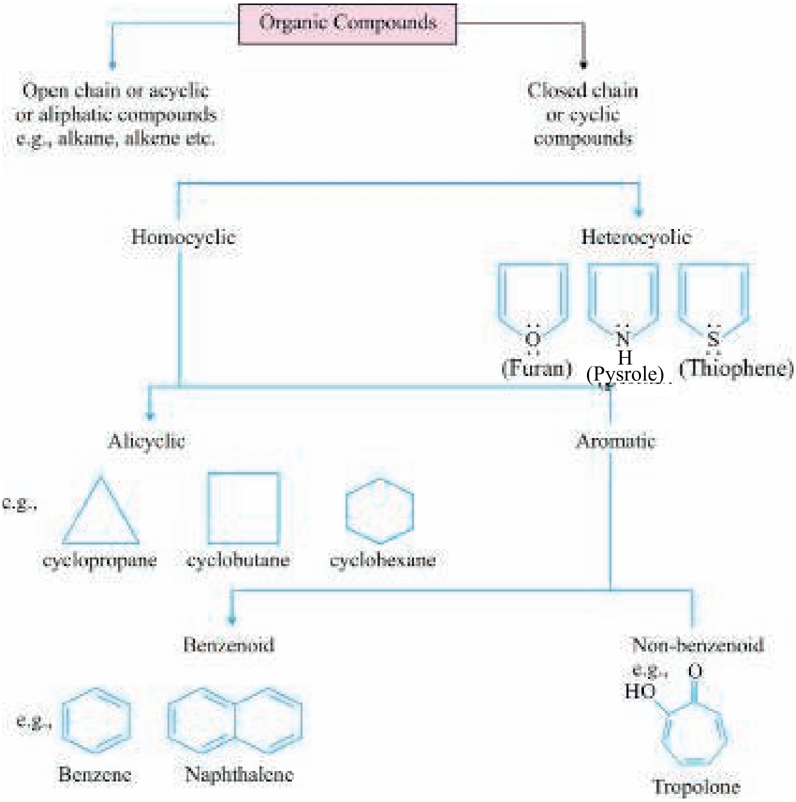
- Acetic acid was synthesised by Kolbe and methane by Berthelot.
- **Types of hybridisation of C-atom :**

sp^3	Tetrahedral	109°28'	Ethane, Methane
sp^2	Trigonal	120°	Ethene, Propene
sp	Linear	180°	Ethyne, Propyne

- **Reasons for existence of large number of organic compounds:**
- **Catenation :** The property of atoms of an element to link with one another forming chains of identical atoms is called *catenation*. Carbon exhibits catenation to the maximum extent.
- **Isomerism :** It is the property by virtue of which two or more compounds have the same molecular formula but different physical or chemical properties.

- **Formation of multiple bonds :** Because of its small size carbon atom is capable of forming multiple bonds with other atoms and this gives a variety of compounds.

- **CLASSIFICATION OF ORGANIC COMPOUNDS**



- **CLASSIFICATION OF CARBON ATOMS**

  On the basis of number of C attached

  (i) **Primary carbon atom :** when carbon atom is attached with one other carbon atom only, it is called **primary or 1°** carbon atom.

  (ii) **Secondary carbon atom :** When carbon atom is attached with two other carbon atoms, it is called **secondary or 2°** carbon atom.

  (iii) **Tertiary carbon atom :** When carbon atom is attached with three other carbon atoms, it is called **tertiary or 3°** carbon atom.

  (iv) **Quaternary carbon atom :** When carbon atom is attached with four other carbon atoms, it is called **Quarternary or 4°** carbon atom.

- **Functional Group :** The atom *e.g.,* –Cl, –Br, etc., or group of atoms *e.g.,* –COOH, –CHO, which is responsible for the chemical properties of the molecule, is called **functional group**.

- **Homologous Series :** The series in which the molecular formula of adjacent members differ by a $-CH_2$ unit, is called homologous series and the individual members are called homologous, *e.g.,* The homologous series of alkene group is

$$\left.\begin{matrix} C_2H_4 \\ C_3H_6 \\ C_4H_8 \\ C_5H_{10} \end{matrix}\right\} \text{difference of} -CH_2 \text{ unit or 14 unit mass}$$

The general characteristics of this series are :

1. All the homologues contain same functional group. That's why their chemical properties are almost similar.
2. All the members of a series have same general formula, *e.g.,*

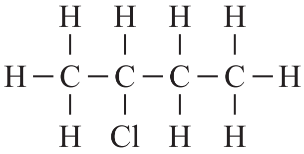
Alkanes	C_nH_{2n+2}
Alkenes	C_nH_{2n}
Alkynes	C_nH_{2n-2}
Alcohol and ether	$C_nH_{2n+2}O$
Aldehyde and ketone	$C_nH_{2n}O$
Acid and ester	$C_nH_{2n}O_2$

3. All the members can be prepared by almost similar methods.
4. With increase in the molecular weight of a series, the physical properties vary gradually.

- **Representation of Organic Compounds :**

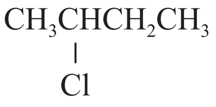
Organic compounds can be represented by the following ways:

(i) **Complete Structural Formula :** All the bonds present between any two atoms are shown clearly. *e.g.,*



$$\begin{matrix} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & Cl & & H & & H & & \end{matrix}$$

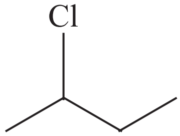
(ii) **Condensed Formula :** All the bonds are not shown clearly. *e.g.*,



$$\underset{\displaystyle Cl}{CH_3\underset{|}{C}HCH_2CH_3}$$

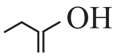
$CH_3CH(Cl)CH_2CH_3$

(iii) **Bond Line Formula :** Every fold and free terminal represents a carbon and lines represent the bond. *e.g.*,



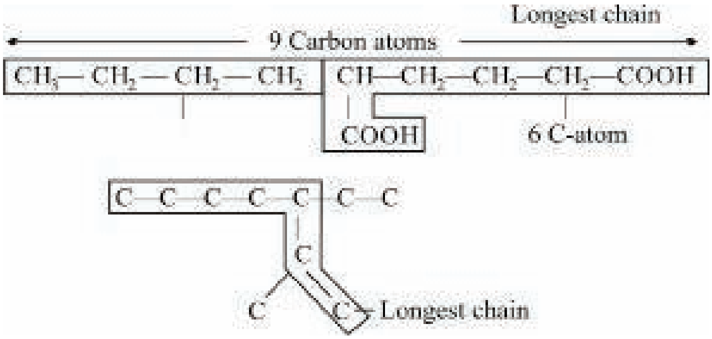
$$CH_3-\underset{\displaystyle CH_3}{\underset{|}{C}}=CH-CH_2CH_3 \equiv$$

$CH_3-CH_2-COOH \equiv$



- **IUPAC Nomenclature of Organic Compounds :** Following rules are used to write the IUPAC name of an organic compound.

Rule 1. : Longest chain rule : The chain containing the principal functional group, secondary functional group and multiple bonds as many as possible is the longest possible chain.
In the absence of functional group, secondary group and multiple bonds, the chain containing the maximum number of C-atoms will be the longest possible chain *e.g.*,



Word Root for Carbon Chain

Chain length	Word root	Chain length	Word root
C_1	Meth-	C_7	Hept
C_2	Eth-	C_8	Oct
C_3	Prop-	C_9	Non
C_4	But-	C_{10}	Dec
C_5	Pent-	C_{11}	Undec
C_6	Hex-	C_{12}	Dodec

Rule 2 : Lowest number rule : Numbering is done in such a way so that

(i) branching if present gets the lowest number.

(ii) the sum of numbers of side chain is lowest.

(iii) principal functional group gets the lowest number.

Select the principal functional group from the preference series :

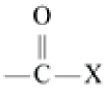
$-COOH > -SO_3H >$ acid anhydride $-COOR > -COX > -CONH_2$
$> -CN > -NC > -CHO > C = O > -OH > -SH$
$> -NH_2 > = > \equiv >$ substituents

Functional group other than the principal functional group are called substituents.

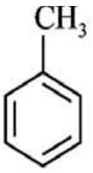
Rule 3: Naming the prefix and suffixes : Prefix represents the substituent and suffix is used for principal functional group.

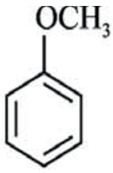
Primary suffix are **ene, ane** or **yne** used for double, single and triple bonds respectively.

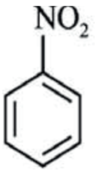
Secondary suffixes are tabulated below :

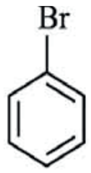
No.	Class	Formula	Prefix	Suffix
1.	Acid halides	 $-\overset{O}{\overset{\|}{C}}-X$	halocarbonyl	—oyl halide —carbonyl halide
2.	Alcohols	—OH	hydroxy	—ol
3.	Aldehydes	—CHO	formyl	—al —carbaldehyde
4.	Ketones	$>C = O$	oxo (keto)	—one
5.	Amides	$-CONH_2$	carbamoyl	—amide
6.	Amine	$-NH_2$	amino	—amine
7.	Carboxylic acid	—COOH	carboxy	—carboxylic acid
8.	Ester	—COOR	alkoxy carbonyl	—alkyl alkan**oate**
9.	Nitriles	—CN	cyano	—nitrile
10.	Sulphonic acid	$-SO_3H$	sulpho	—sulphonic acid

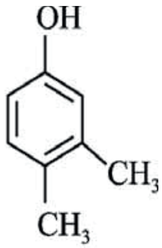
- **Nomenclature of substituted benzene compounds :**

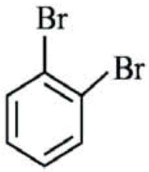

Methylbenzene

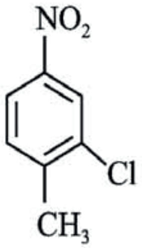

Methoxy benzene


Nitrobenzene


Bremobenzene


3, 4-Dimethyl phenol


1, 2-Dibromo benzene
In Trivial name
1, 2-ortho (o)
1, 3-meta (m)
1, 4-para (p)


2-Chloro-1-methyl 1-4-nitrobenzene

ISOMERISM

Structural isomerism

Same molecular formula but different structures

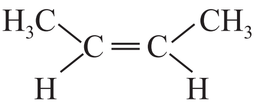
Types

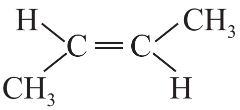
1. **Chain Isomerism**
   e.g., Pentane and
   2-Methylbutane
2. **Position Isomerism**
   e.g., But-1-ene
   But-2-ene
3. **Functional Isomerism**
   e.g., Propanal & Propanone
   Ethanol & Methoxymethane
4. **Metamerism**
   e.g., Pentan-2-one and
   Pentan-3-one

Stereo isomerism

Same molecular and structural formula but different configuration

e.g., Geometrical isomerism (cis-trans isomerism)


Cis.But-2-ene


Trans.But-2-ene

- **Fission of a Covalent Bond :**

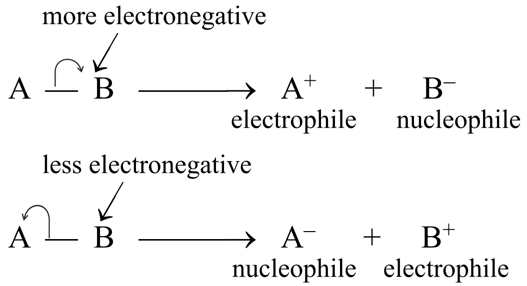
  (i) **Homolytic Fission :** In this one of the electrons of the shared pair in a covalent bond goes with each of the bonded atoms. The neutral chemical species thus formed, are called free radicals. Generally, homolytic fission takes place in non-polar covalent molecules in the presence of sunlight or high temperature.

$$A\overset{\curvearrowright}{—}B \xrightarrow{\text{Sunlight}} \underbrace{A^{*} + B^{*}}_{\text{free radicals}}$$

$$Cl_2 \xrightarrow{\text{Sunlight}} 2Cl^{*}$$

  Free radicals are highly reactive, neutral and electron deficient species.

  (ii) **Heterolytic Fission :** The covalent bond breaks in such a fashion that the shared pair of electrons goes with one of the fragments.



  Heterolytic fission generally takes place in polar covalent molecules but in non-polar molecules, it takes place in the presence of catalyst like $AlCl_3$ (anhy.), $FeCl_3$ (anhy.) etc.

- **Attacking Reagents :**

  These are of two types

  (i) **Electrophiles or Electrophilic Reagents**

  These are electron deficient species, i.e., behave as Lewis acids.

  e.g., Cl^+, $\overset{+}{N}O_2$, CH_3CO^+ etc.

  BF_3, $ZnCl_2$ (anhydrous), $FeCl_3$ (anhydrous), $AlCl_3$ (anhydrous)

  (ii) **Nucleophiles or Nucleophilic Reagents**

  These are negatively charged or neutral molecules with unshared electron pair.

  e.g., $O\bar{H}$, CN^-, $R\overset{\circ\circ}{N}H_2$, $\overset{\circ\circ}{N}H_3$

- **Reaction Intermediates :**

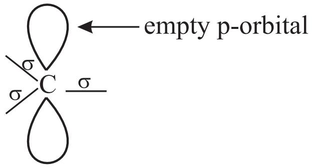
  (i) **Free radicals :** These are the product of homolysis and contain an odd electron. These are highly reactive planar species with sp^2 hybridisation.

Their order of stability is

$$(C_6H_5)_3\overset{\circ}{C} > (C_6H_5)_2\,\overset{\circ}{C}H > C_6H_5\overset{\circ}{C}H_2$$
$$> CH_2 = CH - \overset{\circ}{C}H_2 > 3° > 2° > 1° > CH_2 = \overset{\circ}{C}H$$

(ii) **Carbocations :** These are the product of heterolysis and contain a carbon bearing positive charge. These are electron deficient species.

These are also polar chemical species i.e., sp^2 hybridised with an empty p-orbital.



Stability order of carbocation is

$$(C_6H_5)_3C^+ > (C_6H_5)_2\,\overset{+}{C}H > 3° > C_6H_5\overset{\oplus}{C}H_2 > 2° > CH_2{=}CH_2{-}\overset{+}{C}H_2$$
$$> 1° > CH_2 = \overset{\oplus}{C}H$$

(iii) **Carbanions :** These are the product of heterolysis and contain a carbon bearing negative charge and 8 electrons in its valence shell.

These have pyramidal shape with sp^3 hybridised carbon (having one lone pair) order of stability of carbanions is

$$(C_6H_5)_3\overline{C} > (C_6H_5)_2\,\overline{C}H > C_6H_5\overline{C}H_2$$
$$> \overline{C}H_3 > 1° > 2° > 3° \text{ carbanions}$$

- **Electron Displacement in Covalent Bond**

1. Inductive Effect : If shared pair is more shifted towards more electronegative atom, the less electronegative atom acquires slight positive charge and more electronegative atom acquires partial negative charge,

e.g., $\overset{+\delta}{CH_3} \rightarrow\!\!- \overset{-\delta}{Cl}$

Permanent effect and propagates through carbon chain.

Atoms or groups having greater electron affinity than hydrogen are said to have electron attracting or negative inductive effect (–*I*) while that having, smaller electron affinity than hydrogen are said to have electron releasing or positive inductive effect (+*I*).

e.g., $\overset{+\delta\delta}{CH_3} \rightarrow\!\!- \overset{+\delta}{CH_2} \rightarrow\!\!- \overset{-\delta}{Cl}$

$$\overset{+\delta\delta\delta}{CH_3} \rightarrow\!\!- \underset{1°\text{ alkyl halide}}{\overset{+\delta\delta}{CH_2} \rightarrow\!\!- \overset{+\delta}{CH_2}} \rightarrow\!\!- \overset{-\delta}{Cl}$$

Cl has –I effect and alkyl group has +I effect.

Order of groups producing –I effect is

$R_3N > -NO_2 > -CN > -SO_3H > -CHO > -CO > -COOH > -F$
$> -Cl > -Br > -I > -OH > -OR > -NH_2 > -C_6H_5 > -H$

Order of groups producing +I effect is

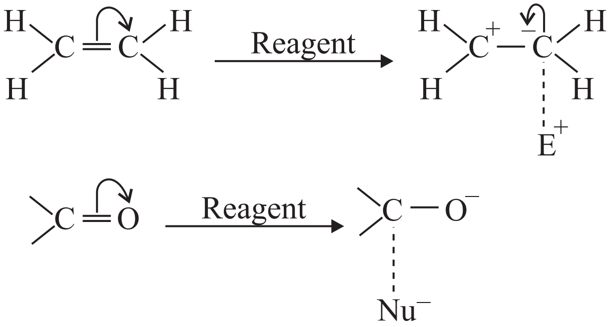
3° alkyl group > 2° alkyl group > 1° alkyl group $> -CH_3 > -H$

- **Applications of Inductive Effect**
  - (i) Presence of groups showing +I effect increases the stability of carbocation while presence of groups showing –I effect decreases their stability.
  - (ii) Strength of acid increases with the attachment of group showing –I effect and decreases with the attachment of group showing +I effect.
  - (iii) Presence of +I showing groups increases the basic strength of amines.

2. **Electromeric Effect :** Defined as the polarity produced in a multiple bonded compound as a reagent approaches it. In the presence of attacking reagent, the two π electrons are completely transferred to any of the one atom. This effect is temporary.

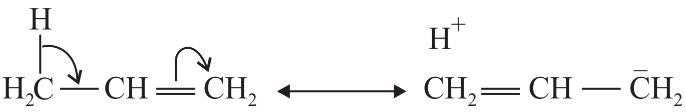
It may be of +E type (when displacement of electron pair is away from the atom of group) or of –E type (when displacement is towards the atom or group).

e.g.,



3. **Hyper-conjugation :** It involves delocalisation of σ electron of a C – H bond of an alkyl group attached directly to an atom of unsaturated system or to an atom with an unshared p-orbital.

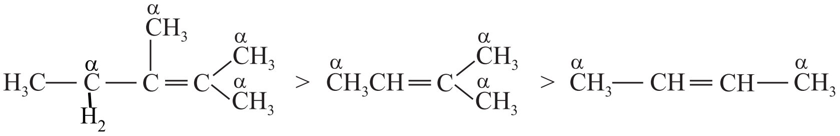
e.g.,



This effect is also called no bond resonance or Baker Nathan effect.

Applications of Hyper-conjugation

Stability of alkenes : More the number of α-hydrogen atoms, more stable is the alkene.



Stability of Carbocation : Greater the number of alkyl groups attached to positively charged carbon atom, the greater is the stability.

e.g., $(CH_3)_2\overset{+}{C}H > CH_3 - \overset{+}{C}H > \overset{+}{C}H_3$

4. **Resonance Effect :** When the properties of a molecule cannot be shown by a single structure and two or more structures are required to show all the properties of that molecule, then the structures are called resonating structures or canonical forms and the molecule is referred as resonance hybrid. This phenomenon is called resonance.

Conditions for resonance

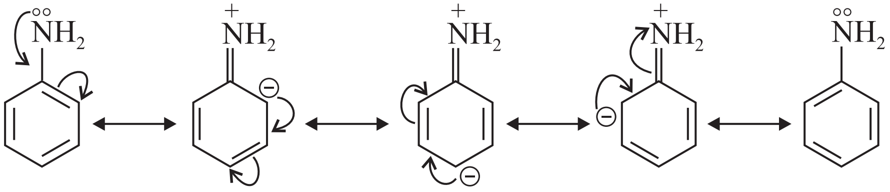
(i) The arrangement of atoms must be identical in all the formula.

(ii) The energy content of all the canonical forms must be nearly same.

(iii) Each canonical of π electrons. This effect may be of +R type or –R type.

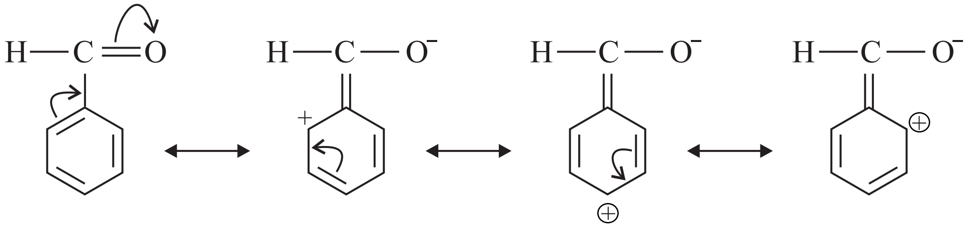
Positive Resonance Effect (+R)

Electron donating groups with respect to conjugated system show +R effect. Central atom of functional groups should be more electronegative than the surrounding atoms or groups to show +R effect. *e.g.*, halogens, –OH, –OR, $-NH_2$, NHCOR, etc.



Negative Resonance Effect (–R)

Electron withdrawing groups with respect to conjugate system show –R effect. Central atom of functional groups should be less electronegative than surrounding atoms or groups to show –R effect. *e.g.*, halogens, –COOH, –COOR, CHO, –CN, $-NO_2$, etc.



Methods of Purification of Organic Compounds

Method	Principle	Applications
Crystallization	Different solubilities of a given organic compound and its impurities in the same solvent.	- Crystallization of sugar (containing an impurity of common salt) is achieved by shaking the impure solid with hot ethanol at 348K (sugar dissolves whereas common salt remains insoluble).
Sublimation	Some solid substances change from solid to vapour state without passing through liquid state. Sublimable compounds get separated from non-sublimable impurities.	- Iodine from sodium chloride (as iodine sublimes readily leaving behind sodium chloride). - Camphor, naphtalene, anthracence, benzoic Acid, etc. are purified.
Distillation	It is used to separate - Volatile liquids from non-volatile impurities. - Liquids having sufficient difference in their boiling points.	- Hexane (b.p. 342K) and toluene (b.p. 384K) - Chloroform (b.p. 334K) and aniline (b.p. 457K)
– Fractional Distillation	If the difference in boiling points of two liquids is not much, this method is used.	- Crude oil in petroleum industry is separated into various useful fractions such as gasoline, kerosene oil, diesel oil, lubricating oil, etc.
– Steam Distillation	This method is used to separate substances which are (i) steam volatile, (ii) immiscible with water, (iii) posses a vapour pressure of 10-15 mm Hg and (iv) contain non-volatile impurities.	- Aniline is separated from aniline water mixture. - Essential oils, turpentine oil, o-nitrophenol, bromobenzene nitrobenenze, etc. can be purified.
Differential Extraction	By shaking an aqueous solution of an organic compound with an organic solvent in which the organic compound is more soluble than in water. The organic solvent and the aqueous solution should be immiscible with each other so that they can form two distinct layers which can be separated by using separating funnel.	- Benzoic acid can be extracted from its water solution using benzene.

Chromatography	Differential movement of individual components of a mixture through a stationary phase under the influence of a mobile phase.	• Widely used for separation purification, identification and characterization of the components of a mixture, whether coloured or colourless.
– Adsorption Chromatography	Differential adsorption of the various components of a mixture on a suitable adsorbent such as silica get or alumina.	
– Column Chromatography	The mixture is passed through adsorbent packed in glass tube.	• Mixture of naphthalene and benzophenone.
– Thin Layer Chromatography	The mixture is passed over adsorbent on a thin glass plate.	• Amino acids can be detected by spraying the plate with ninhydrin solution.
– Partition Chromatography	Differential partitioning of components of a mixture between stationary and mobile phases.	
– Paper Chromatography	A special quality paper known as chromatography paper is used. It contains water trapped in it, which acts as the stationary phase.	• For separation of sugars and amino acids.

Types of Chromatography	Mobile / Stationary Phase
Column Chromatography	Liquid / Solid
Thin Layer Chromatography	Liquid / Solid
High Performance Liquid Chromatography (HPLC)	Liquid / Solid
Gas Liquid Chromatography (GLC)	Gas / Solid
Partition or Paper Chromatography	Liquid / Solid

Element	Detection	Confirmatory Test	Reactions
Carbon	$2CuO + C \xrightarrow{\Delta} 2Cu + CO_2$	CO_2 gas turns lime water milky.	$CO_2 + \underset{\text{Lime water}}{Ca(OH)_2} \longrightarrow \underset{\text{Milkiness}}{CaCO_3\downarrow} + H_2O$
Hydrogen	$CuO + 2H \xrightarrow{\Delta} Cu + H_2O$	Water droplets appear on the cooler part of the ignition tube and also turns anhydrous $CuSO_4$ blue.	$\underset{\text{White}}{CuSO_4} + 5H_2O \longrightarrow \underset{\text{Blue}}{CuSO_4.5H_2O}$
Nitrogen	Lassaigne's extract (L.E.) $Na + C + N \xrightarrow{\Delta} \underset{\text{(L.E.)}}{NaCN}$	L.E. + $FeSO_4$ + NaOH, boil and cool + $FeCl_3$ + conc. HCl. Gives blue or green colour.	$FeSO_4 + 2NaOH \longrightarrow Fe(OH)_2 + Na_2SO_4$ $Fe(OH)_2 + 6NaCN \longrightarrow Na_4[Fe(CN)_6] + 2NaOH$ $3Na_4[Fe(CN)_6] + 4FeCl_3 \longrightarrow \underset{\text{Prussian blue}}{Fe_4[Fe(CN)_6]_3} + 12NaCl$

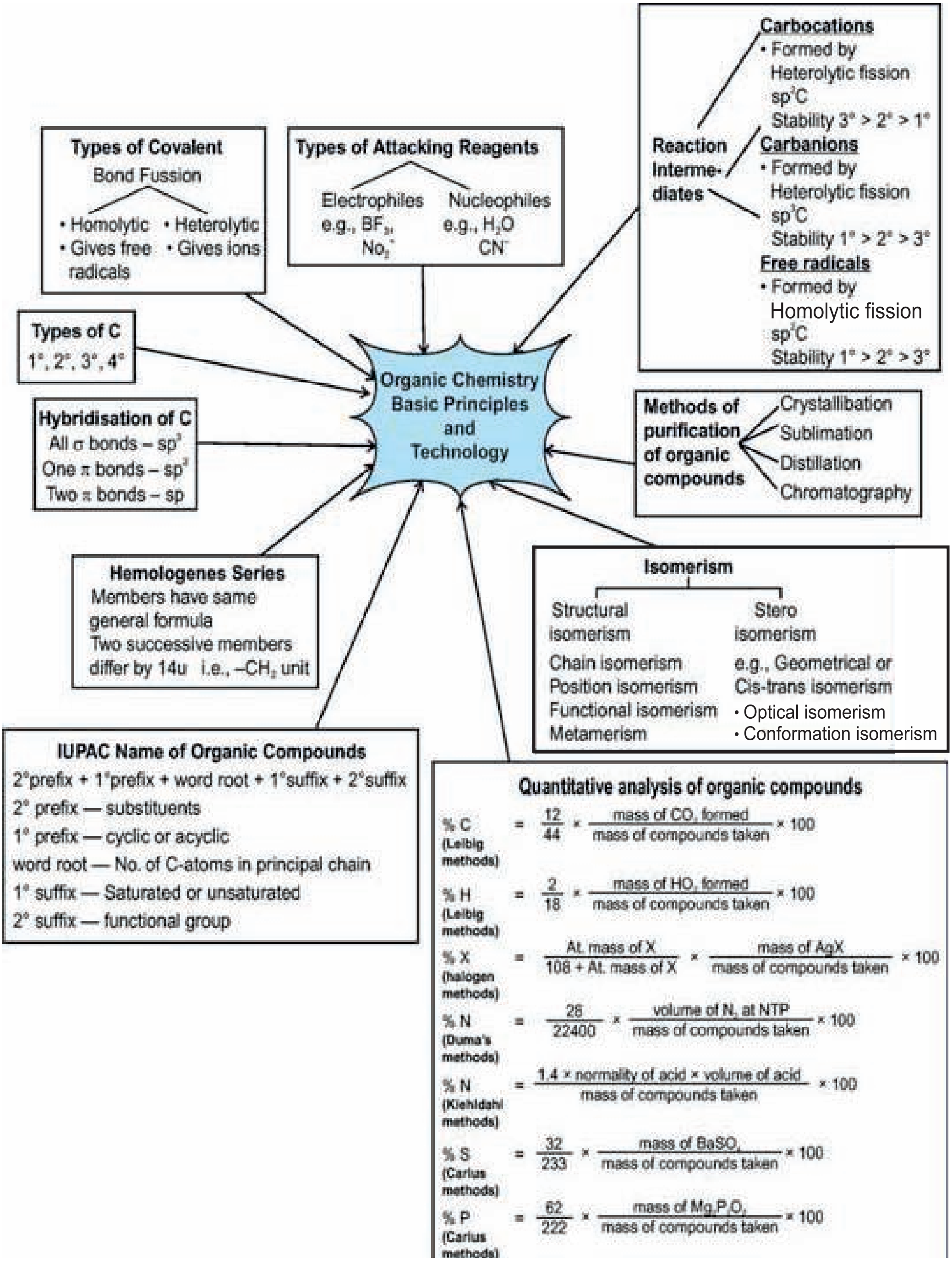
Sulphur	$2Na + S \xrightarrow{\Delta} Na_2S$ (L.E.)	– L.E. + sodium nitroprusside A deep violet colour. – L.E. + CH_3COOH + $(CH_3COO)_2Pb$ Gives a black ppt.	$Na_2S + Na_2[Fe(CN)_5NO] \longrightarrow$ Sodium nitroprusside $Na_4[Fe(CN)_5NOS]$ Deep violet $Na_2S + (CH_3COO)_2Pb \xrightarrow{CH_3COOH}$ $Pbs\downarrow + 2CH_3COONa$ Black ppt.
Halogens	$Na + X \xrightarrow{\Delta} NaX$ (L.E.)	$L.E. + HNO_3 + AgNO_3$ – White ppt. soluble in aq. NH_3 (or NH_4OH) confirms Cl. – Yellow ppt. partially soluble in aq. NH_3 (or NH_4OH) confirms Br. – Yellow ppt. insoluble in aq. NH_3 (or NH_4OH) confirms I.	$NaX + AgNO_3 \xrightarrow{HNO_3} AgX\downarrow$ ppt. $AgCl + 2NH_{3(aq.)} \longrightarrow [Ag(NH_3)_2]Cl$ White ppt. Soluble
Nitrogen and sulphur together	$Na + C + N + S \xrightarrow{\Delta}$ $NaSCN$ Sodium thiocyanate (L.E.)	As in test for nitrogen; instead of green or blue colour, blood red colouration confirms presence of N and S both.	$NaSCN + FeCl_3 \longrightarrow$ $[Fe(SCN)Cl_2] + NaCl$ Blood red colour
Phosphorus	$P \xrightarrow{Na_2O_2,\ boil} Na_3PO_4$	Solution is boiled with nitric acid and then treated with ammonium molybdate $(NH_4)_2MoO_4$. Formation of yellow ppt. indicates presence of phosphate (hence, phosphorus) in organic compound.	$Na_3PO_4 + 3HNO \longrightarrow H_3PO_4 + 3NaNo_3$ $H_3PO_4 + 12(NH_4)_2MoO_4 + 21HNO_3$ $\longrightarrow (NH_4)_3PO_4.12MoO_3$ Ammonium phosphomolybalate (yellow ppt.) $+ 21NH_4NO_3 + 12H_2O$

- **Quantitative analysis of organic compounds :** The percentage composition of elements presence an organic compound is determined by the methods based on the following principles :

Elements	**Method**
Carbon and Hydrogen	**Liebig's Combustion method :** A known mass of an organic compound is burnt in the presence of excess of O_2 and CuO. $C_xH_y + \left(x + \frac{y}{4}\right)O_2 \xrightarrow{\Delta} xCO_2 + \frac{y}{2}H_2O$ CO_2 evolved is absorbed by conc. solution of KOH or ascarite (NaOH + CaO).H_2O produced is absorbed by anhydrous $CaCl_2$ or $Mg(ClO_4)_2$. Increase in masses of these absorbing compounds gives the masses of CO_2 and H_2O produced. $\% \text{ of C} = \frac{12}{44} \times \frac{\text{mass of } CO_2 \text{ formed}}{\text{mass of compound taken}} \times 100;$ $\% \text{ of H} = \frac{2}{18} \times \frac{\text{mass of } H_2O \text{ formed}}{\text{mass of compound taken}} \times 100$
Halogens	**Carius method :** Halogen in organic compound is precipitated as silver halide by boiling with conc. NHO_3 and then adding $AgNO_3$. $X \xrightarrow[AgNO_3]{HNO_3,\ \Delta} AgX\downarrow$ $\% \text{ of X} = \frac{\text{At. mass of X}}{108 + \text{At. mass of X}} \times \frac{\text{mass of AgX formed}}{\text{mass of compound taken}} \times 100$
Nitrogen	**Dumas method :** Nitrogen containing organic compound is heated with CuO in an atmosphere of CO_2. $C_xH_yN_z + \left(2x + \frac{y}{2}\right)CO_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O + N_2 + \left(2x + \frac{y}{2}\right)Cu$ N_2 evolved gets collected over conc. KOH solution which absorbs all other gases. $\% \text{ of N} = \frac{28}{22400} \times \frac{\text{Vol. of } N_2 \text{ at STP}}{\text{mass of compound taken}} \times 100$ **Kjeldahl's method :** Organic compound + $H_2SO_{4\,(\text{conc.})} \longrightarrow (NH_4)_2SO_4 \xrightarrow{2NaOH} Na_2SO_4 +$ $2NH_3 + 2H_2O + 2NH_3 + H_2SO_4 \longrightarrow (Na_4)_2SO_4$ $\% \text{ of N} = \frac{1.4 \times \text{molarity of acid} \times \text{vol. of acid used} \times \text{basicity of acid}}{\text{mass of compound taken}}$

Sulphur	**Carius method :** Sulphur in organic compound is converted into H_2SO_4 by boiling with Na_2O_2 or conc. HNO_3 and is precipitated as $BaSO_4$ by adding excess of $BaCl_2$ solution in water. $S \xrightarrow[\text{(ii) BaCl}_2]{\text{(i) HNO}_3\text{, }\Delta} \underset{\text{white ppt.}}{BaSO_4\downarrow}$ $\% \text{ of S} = \frac{32}{233} \times \frac{\text{mass of BaSO}_4\text{ formed}}{\text{mass of compound taken}} \times 100$
Phosphorus	**Ignition method :** $P \xrightarrow[\text{heat}]{HNO_3} H_3PO_4$ $H_3PO_4 + Mg^{2+} + NH_4CL \xrightarrow{\Delta} \underset{\text{Magnesium ammonium phosphate (white ppt.)}}{MgNH_4PO_4} + HCl$ $2MgNH_4PO_4 \xrightarrow{\Delta} \underset{\text{Magnesium pyrophosphate}}{Mg_2P_2O_7} + 2NH_3 + H_2O$ $\% \text{ of P} = \frac{62}{222} \times \frac{\text{mass of Mg P}_2\text{O}_7\text{ formed}}{\text{mass of compound taken}} \times 100$

MIND MAP : ORGANIC CHEMISTRY



Organic Chemistry Basic Principles and Technology

Types of Covalent
Bond Fussion
- Homolytic
- Gives free radicals
- Heterolytic
- Gives ions

Types of Attacking Reagents
Electrophiles e.g., BF_3, NO_2^+
Nucleophiles e.g., H_2O, CN^-

Reaction Interme-diates

Carbocations
- Formed by Heterolytic fission
sp^2C
Stability 3° > 2° > 1°

Carbanions
- Formed by Heterolytic fission
sp^3C
Stability 1° > 2° > 3°

Free radicals
- Formed by Homolytic fission
sp^2C
Stability 1° > 2° > 3°

Types of C
1°, 2°, 3°, 4°

Hybridisation of C
All σ bonds – sp^3
One π bonds – sp^2
Two π bonds – sp

Methods of purification of organic compounds
- Crystallibation
- Sublimation
- Distillation
- Chromatography

Hemologenes Series
Members have same general formula
Two successive members differ by 14u i.e., $-CH_2$ unit

Isomerism

Structural isomerism
- Chain isomerism
- Position isomerism
- Functional isomerism
- Metamerism

Stero isomerism
- e.g., Geometrical or Cis-trans isomerism
- Optical isomerism
- Conformation isomerism

IUPAC Name of Organic Compounds
2°prefix + 1°prefix + word root + 1°suffix + 2°suffix
2° prefix — substituents
1° prefix — cyclic or acyclic
word root — No. of C-atoms in principal chain
1° suffix — Saturated or unsaturated
2° suffix — functional group

Quantitative analysis of organic compounds

% C (Leibig methods) $= \frac{12}{44} \times \frac{\text{mass of } CO_2 \text{ formed}}{\text{mass of compounds taken}} \times 100$

% H (Leibig methods) $= \frac{2}{18} \times \frac{\text{mass of } H_2O \text{ formed}}{\text{mass of compounds taken}} \times 100$

% X (halogen methods) $= \frac{\text{At. mass of X}}{108 + \text{At. mass of X}} \times \frac{\text{mass of AgX}}{\text{mass of compounds taken}} \times 100$

% N (Duma's methods) $= \frac{28}{22400} \times \frac{\text{volume of } N_2 \text{ at NTP}}{\text{mass of compounds taken}} \times 100$

% N (Kjehldahl methods) $= \frac{1.4 \times \text{normality of acid} \times \text{volume of acid}}{\text{mass of compounds taken}} \times 100$

% S (Carius methods) $= \frac{32}{233} \times \frac{\text{mass of } BaSO_4}{\text{mass of compounds taken}} \times 100$

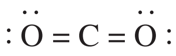
% P (Carius methods) $= \frac{62}{222} \times \frac{\text{mass of } Mg_2P_2O_7}{\text{mass of compounds taken}} \times 100$

CASE BASED STUDY - QUESTIONS

PASSAGE -I

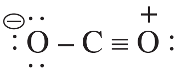
The resonance effect is defined as the polarity produced in the molecule by the interaction of two π bonds as between a π bond and lone pair of electrons present an on adjacent atom. The effect is transmitted through the chain. In positive resonance effect the transfer of electrons is away from on atom or subsitituent group attached to the conjugated system. The electron displacement makes certain positions in the molecule of high electron densities. In negative resonance effect the transfer of electrons is towards the atom or substituent group attached to the conjugated system.

1. Draw resonance structure of carboxylate ion
2. What do you understand by +R and –R effect.
3. Write resonance structures of $CH_2=CH-CHO$ indicate relative stability of the contributing structures.
4. Draw the resonating structures of (a) phenol (b) Benzoic acid
5. Out of the resonating structures of CO_2 which structures is not correct for CO_2?



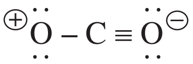
$:\ddot{O}=C=\ddot{O}:$

(i)



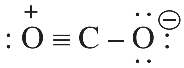
$^{\ominus}:\underset{\cdot\cdot}{\ddot{O}}-C\equiv\overset{+}{O}:$

(ii)



$^{\oplus}\underset{\cdot\cdot}{\ddot{O}}-C\equiv\underset{\cdot\cdot}{\ddot{O}}^{\ominus}$

(iii)



$:\overset{+}{O}\equiv C-\underset{\cdot\cdot}{\ddot{O}}:^{\ominus}$

(iv)

PASSAGE -II

The Lassaigne's extract is usually alkaline because excess of sodium reacts with water to form sodium hydroxide. If not it may be made alkaline by the addition of a few drops of a dilute solution of sodium hydroxide. To a part of the extract a small amount of a freshly prepared ferrous sulphate solution is added and the content are warmed. A few drops of ferric chloride solution are then added to the contents and the resulting solution is acidified with dilute hydrochloric acid the appearance of a bluish green colour due to the formation of ferric ferrocyanide confirm the presence of nitrogen in the organic compound.

1. In sodium fusion test of organic compound the nitrogen in an organic compound is converted into which compound.
2. In the Lassaigne's test for the detection of sulphure the purple colour is due to the formation of which complex.
3. What is the name of the Prussian blue colour compound formed in Lassaigne's test for nitrogen in an organic compound.
4. Why do we boil Lassaigne's extract with conc. HNO_3 while detecting halogens in an organic compound.
5. Explain the reason for the fusion of an organic compound with metallic sodium for testing nitrogen, sulphure and halogens.

MULTIPLE CHOICE QUESTIONS (MCQ)

1. Homolytic fission of C-C bond in ethane gives an intermediate in which carbon is:
   a. sp^3 hybridised
   b. sp^2 hybridised
   c. sp-hybridised
   d. sp^3d- hybridized

2. The kind of delocalization involving sigma bond in conjugation with pi electrons is called:
   a. Inductive effect
   b. Hyperconjugation effect
   c. Electromeric effect
   d. Mesomeric effect

3. Which organic species has only one type of hybridized carbon?
   a. $CH_2{=}C{=}CH_2$
   b. $CH_3{-}\overset{\oplus}{C}H{-}CH_3$
   c. $CH_3{-}C{=}CH$
   d. $CH_2{=}CH{-}\overset{\oplus}{C}H_2$

4. Which of the following can act as an electrophile?
   a. CN^-
   b. OH^-
   c. H_2O
   d. BF_3

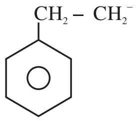
5. Which of the following is correct about the species: $(CH_3)_3{-}C^+$
   a. It is planar
   b. Its C+ is sp^2 hybridised
   c. A nucleophile can attack on its C^+
   d. All of these

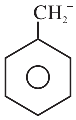
6. Which of the following has all the effects namely Inductive, Mesomeric and Hyperconjugative ?
   a. CH_3Cl
   b. $CH_3CH{=}CH_2$
   c. $CH_3CH{=}CHCOCH_2Cl$
   d. $CH_2{=}CH{-}CH{=}CH_2$

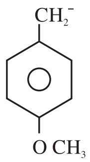
7. The most stable free radical among the following is:
   a. $C_6H_5CH_2\dot{C}H_2$
   b. $C_6H_5\dot{C}HCH_3$
   c. $CH_3\dot{C}H_2$
   d. $CH_3\dot{C}HCH_3$

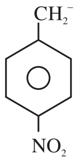
8. Isomers of a compound must have :
   a. Same physical properties
   b. Same chemical properties
   c. Same structural properties
   d. Same molecular weight

9. Most stable carbanion among the following is

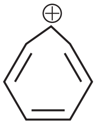
a. 

b. 

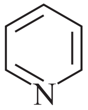
c. 

d. 

10. Which of the following species have six π conjugated electrons?

a. 

b. $CH_2=CH-CH=CH-CH_2^-$

c. 

d. All of these

11. The correct decreasing order of priority for the functional groups of organic compounds in the IUPAC system of nomenclature is:

a. $-COOH, -SO_3H, -CONH_2, -CHO$

b. $-SO_3H, -COOH, -CONH_2, -CHO$

c. $-CHO, -COOH, -SO_3H, -CONH_2$

d. $-CONH_2, -CHO, -SO_3H, -COOH$

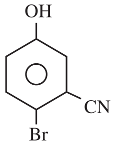
12. The IUPAC name of $CH_3-CH=CH-C\equiv CH$ is:

a. pent-3-en-1-yne

b. pent-3-en-4-yne

c. pent-2-en-4-yne

d. pent-2-en-3-yne

13. The IUPAC name of the following compound is 

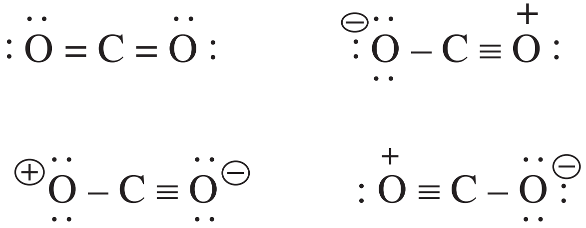
a. 4–Bromo –3– cyanophenol

b. 2–Bromo–5–hydroxy benzonitrile

c. 2–cyano–4–hydroxybromo benzene

d. 6–Bromo–3–hydroxy benzonitrile

14. The order of the stability of the following of carbocations is :



a. III > I > II

b. III > II > I

c. II > III > I

d. I > II > III

15. Quantitative measurement of nitrogen in an organic compound is done by the method:

a. Berthelot method

b. Lassaigne method

c. Carius method

d. Kjehldahl method

ANSWERS

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

11. a 12. a 13. b 14. a 15. d

FILL IN THE BLANKS

1. A triple bond between two carbon atoms is composed of one _________ and _________ bonds.
2. An organic compound which decomposes below its boiling point can be purified by ____________.
3. Electrophiles are the species which attack the regions of ___ electron density.
4. Hyperconjugation effect is also known as ___________ resonance.
5. In Duma's method, the nitrogen present in an organic compound is set free as ____________.

ANSWERS

1. σ , two π 2. Vacuum distillation 3. high 4. No-bond 5. Nitrogen

TRUE AND FALSE TYPE QUESTIONS

1. Ethanol and methoxymethane are position isomers.
2. A free radical is a species with an unpaired valence electron.
3. Inductive effect is observed in π bond in presence of attacking reagent.
4. The percentage of carbon and hydrogen are estimated simultaneously in an organic compound by Liebig method.
5. Chromatography is the method used to separate and purify compounds when present in small amounts.

ANSWERS

1. F 2. T 3. F 4. T 5. T

ASSERTION REASON TYPE QUESTIONS

The questions given below are Assertion (A) and Reason (R). Use the following key to select the correct answer.

(a) If both assertion and reason are correct and reason is correct explanation for assertion.

(b) If both assertion and reason are correct but reason is not correct explanation for assertion.

(c) If reason is correct but assertion is incorrect.

(d) If both assertion and reason are incorrect.

1. Assertion: But-1-ene and 2-Methylprop-1-ene are position isomers.
   Reason: Position isomers have same molecular formula but different arrangement of carbon atoms.

2. Assertion: Duma's method is more applicable to nitrogen containing organic compounds than the Kjeldahl's method.
   Reason: Kjeldahl's method does not give satisfactory result in which nitrogen atom is directly attached to oxygen atom.
3. Assertion: Alkanes having more than three carbon atoms exhibit chain isomerism.
   Reason: All carbon atoms in alkanes are sp-hybridised.
4. Assertion: In $CH_2=C=CH_2$, all the carbon atoms are sp^2 hybridised.
   Reason: All the hydrogen atoms lie in one plane.
5. Assertion: Butane and 2-Methylbutane are homologues.
   Reason: Butane is a straight chain alkane while 2-Methylbutane is branched chain alkane.
6. Assertion: Tertiary carbocations are generally formed more easily than primary carbocations.
   Reason: Hyperconjugation as well as inductive effect due to additional alkyl groups stabilize tertiary carbocations.
7. Assertion: Alkyl carbanions like ammonia have pyramidal shape.
   Reason: The carbon atom carrying negative charge has an octet of electrons
8. Assertion: Carbocations are planar in nature.
   Reason: Carbocations are sp^2 Hybridised.
9. Assertion: IUPAC name of compound $CH_3CH=CH-CHO$ is But-2-enal.
   Reason: Functional group gets preference over multiple in IUPAC name of a compound.
10. Assertion: Glycerol is purified by distillation under reduce pressure.
    Reason: Organic compounds in liquid state are purified by distillation.

ANSWERS

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

MATCH THE COLUMNS

Match the statements (a,b,c,d) in column I with the statements (I,ii,iii,iv) in column II.

1.

Column I	Column II	Column II
a. Leibig method	i. N_2	e) Aniline
b. Dumas method	ii. AgX	f) Halogens
c. Kjehldahl method	iii. CO_2 and H_2O	g) Schiff's Nitrometer
d. Carius method	iv. NH_3	h) $CaCl_2$ tube

2.

Column I	Column II
a. Nonbenzenoid aromatic compound	i. 50% s character
b. Catenation	ii. Species containing single unpaired nonbonding electrons
c. Free radical	iii. Chain-forming property of an element
d. sp-hybridised carbon atom	iv. Tropolone

ANSWERS

1. a. iiii, h b. i, g c. iv, e d. ii, f
2. a. iv b. iii c. ii d. i

ONE WORD ANSWER TYPE QUESTIONS

1 Name one common adsorbent in column chromatography.

2. Mention the hybridisation of underlined carbon in $CH_3\underline{C}\equiv N$.

3. What type of isomerism is shown by Pentane and 2-Methylbutane?

4. Nucleophiles are Lewis acids or Lewis bases?

5. What type of bond fission results in the formation of free radicals?

6 . What is the number of electrons present in the outermost shell of carbon in the methyl radical?

7. What is the other name for no-bond resonance?

8. What is the name of the Prussian blue coloured compound formed in Lassaigne's test for nitrogen in an organic compound?

9. SO_3 is an electrophile or nucleophile in sulphonation reaction of benzene?

10. Name suitable technique of separation of the components from a mixture of calcium sulphate and camphor.

1-MARK QUESTIONS

1. Which unique property of carbon is responsible for the large number of carbon compounds?
2. Which has the longest C–C bond length among ethane, ethene and ethyne.
3. How many secondary carbon atoms are present in 2-Methylpentane?
4. Draw structure of 3-Isopropyl-2-methylhexane.

5. Draw bond line structure of $CH_3(CH_2)_6CH{=}CH(CH_2)_2{-}COOH$

6. What are the bond angles in sp^3, sp^2 and sp hybrid orbitals?.

7. Write the correct of priority of the following functional groups:

   $-C{\equiv}N$, $>C{=}O$, $-OH$, $-COOH$, $-CONH_2$

8. Write IUPAC name of :

   (i) $CH_3{-}CH_2{-}CN$

   (ii) $CH_2{=}CHCH_2OH$

   (iii) $CH_3CH_2CH(CH_3){-}CO{-}CH_2CH_3$

   (iv) $CH_3CH_2{-}O{-}CH_2CH(CH_3)CH_3$

   (v) $Cl{-}CH_2{-}C{\equiv}CH$

9. What type of isomerism is exhibited by Propanal and Propanone?

10. Classify the following into electrophiles and nucleophiles:

    H+, NH_3, $AlCl_3$, NO_2^+, CN^- , H_2O, ROH, RNH_2, Carbocation

11. What type of attacking reagents are produced by heterolytic cleavage of covalent bond?

12. Name each of the following species and indicate which member of each pair is more stable:

    (i) CH_3^+, $CH_3CH_2^+$

    (ii) $C_6H_5\overset{+}{C}HCH_3$, $CH_3\overset{+}{C}HCH_3$

    (iii) $CH_2{=}CH{-}\dot{C}H_2$, $\dot{C}H{=}CH{-}CH_3$

    (iv) $CH_3{-}\overset{-}{C}H_2$, $CH_3{-}\overset{-}{C}H{-}CH_3$

13. Identify electrophilic centre in CH_3CHO.

14. What is state of hybridization of positively charged carbon atom in carbocation?

15. What is the effect of introducing an alkyl group on the stability of carbocation?

16. Out of Benzyl and ethyl carbocation which is more stable and why?

17. Arrange the following in increasing order of acidic strength:

    $ClCH_2COOH$, CH_3CH_2COOH, $ClCH_2CH_2COOH$

18. Name two solvents which are commonly used to dissolve organic solids.

19. Name the technique that can be used for purification of iodine that contains traces of NaCl.

20. A liquid (10 mL) has three components A, B, C. which technique is suitable to sparate A,B, C from such a small amount of mixture?

21. Under what condition do we use fractional distillation?

22. A liquid compound starts decomposing well before its boiling point under normal pressure. How will you purify it?

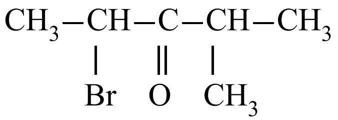
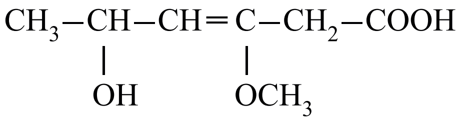
23. For which type of compounds Kjehldahl's method is not useful?

24. How do you precipitate sulphur in Carius method?

25. Which method is used to estimate carbon and hydrogen?

26. What do we notice in Lassaigne's test if the compound contains both nitrogen and sulphur?

2-MARKS QUESTION

1. How will you account for the presence of large number of organic compounds?
2. Draw the structural formulae of the following compounds:
   (i) Ethoxypropane
   (ii) But-1-en-3-yne
   (iii) 3,4,4,-Trimethylhex-1-yne
   (iv) sec-butyl alcohol
   (v) But-2-enoic acid
3. Give IUPAC name of the following compounds:
   (i) $C_6H_5\,CH_2\,CH_2OH$
   (ii) $(CH_3)_2\,CH_2\,CH_2\,CHO$
   (iii) $CH_2{=}CH{-}C{\equiv}N$
   (iv) 
   $CH_3{-}\underset{Br}{\underset{|}{CH}}{-}\underset{O}{\underset{\|}{C}}{-}\underset{CH_3}{\underset{|}{CH}}{-}CH_3$
   (v) 
   $CH_3{-}\underset{OH}{\underset{|}{CH}}{-}CH{=}\underset{OCH_3}{\underset{|}{C}}{-}CH_2{-}COOH$
4. Indicate sigma (σ) and pi (p) bonds in the following molecules

   (i) CH_2Cl_2 (ii) $HCONH.CH_3$

5. Write bond line formulas for: (i) isopropyl alcohol (ii) heptane-4-one

6. Which is suspected to more stable:

   $O_2NCH_2CH_2O^-$ or $CH_3CH_2O^-$ and why?

7. What is the effect of introducing an alkyl group on the stability of a free radical?

8. Give two examples each of the groups exerting –I and + I effect when attached to a chain of carbon atoms.

9. A tertiary butyl carbocation is is more stable than isobutyl carbocation. Justify.

10. All electrophiles are Lewis acids while nucleophile are Lewis bases. Explain.

11. What is the purpose of filtration through hot water funnel?

12. What precautions are necessary while purifying an organic solid with the help of crystallization process?

13. Discuss the principle of steam distillation.

14. Discuss the role of fractionating column in fractional distillation.

15. How will you prepare Lassaigne's extract? Name the elements which can be detected from this extract?

3-MARKS QUESTIONS

1. Why stability of carbocations follows the order: tertiary>secondary>primary?

2. What are the various conditions essential for resonance?

3. Write resonance structures of $CH_2{=}CH{-}CHO$. Indicate relative stability of the contributing structures.

4. Inductive effect is of permanent nature while electromeric effect is only temporary. Explain.

5. What is chromatography? Name different types of chromatographic processes.

6. You are given a mixture of methanol and acetone. Discuss the process which you will employ to separate them.

5-MARK QUESTIONS

1. Classify the following reactions in one of the reaction type studies in this unit:

   (i) $CH_3CH_2Br \longrightarrow CH_3CH_2SH + Br^-$

   (ii) $(CH_3)_2C = CH_2 + HCl \longrightarrow (CH_3)_2C(Cl)CH_3$

   (iii) $(CH_3)_3CCH_2OH + HBr \longrightarrow (CH_3)_2CBrCH_2CH_3 + H_2O$

   (iv) $CH_3CH_2Br + HO^- \longrightarrow CH_2 = CH_2 + H_2O + Br^-$

   (d) $CH_3CH_2Cl + \underset{aq}{KOH} \longrightarrow CH_3CH_2OH + KCl$

2. (a) An organic compound contains 69% carbon and 4.8% hydrogen the remainder being oxygen. Calculate the masses of carbon dioxide and water produce when 0.20 g of this compound is subjected to complete combustion.

   (b) 0.3780 of an organic compound gave 0.574 g of silver chloride in carious estimation. Calculate the percentage of chlorine in the compound.

3. Arrange the following in the order of property indicated against each set:

   (i) $-C_6H_5$, $-NO_2$, $-COOH$, $-I$, $-F$, $-CH_3$, $-C_2H_5$ (In the increasing order of –I effect)

   (ii) $CH_3CH_2CH_2CH_2^+$, $(CH_3)_3C^+$, $CH_3CH_2CH_2\overset{+}{C}HCH_3$ (In the order of increasing stability)

   (iii) $-Cl$, $-CONH_2$, $-CHO$ (In the increasing priority order if present in same molecule)

4. Draw the resonance structures for the following compounds. Show the electron shift using curved arrow notation.

   (i) $C_6H_5NO_2$

   (ii) $CH_3CH{=}CHCHO$

   (iii) C_6H_5OH

   (iv) $C_6H_5CH_2^+$

   (v) $CH_3CH{=}CHCH_2^+$

5. Suggest a method to separate the constituents from the following mixture:
    - (i) Mixture of two miscible liquids
    - (ii) A mixture of oil and water
    - (iii) A mixture of plant pigments
    - (iv) A mixture of solid benzoic acid and sodium chloride
    - (v) o-Nitrophenol and p-Nitrophenol present in the mixture.

6. 0.378g of an organic compound containing carbon and hydrogen was subjected to combustion by Leibig's method, the CO_2 and H_2O formed were passed through potash bulbs and anhydrous $CaCl_2$ tube. At the end of the experiment, the increase in the respective weights were 0.264g and 0.162g. Calculate the percentage of carbon and hydrogen.

    (**Ans:** C = 19.05% , H = 4.76%)

UNIT TEST-I

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

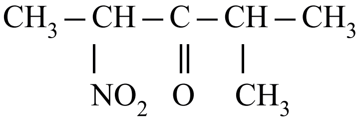
(ii) Maximum marks carried by each question are indicated against it.

---

1. Write bond line formula for the following compound: [1]

   $HOCH_2CH_2CH_2CH(CH_3)CH(CH_3)COOH$

2. Write IUPAC name of the following compound: [1]



$$\underset{\displaystyle NO_2}{CH_3-\underset{|}{CH}}-\underset{\displaystyle O}{\underset{\|}{C}}-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3$$

3. The central atom of compound $CH_2=C=CH_2$ is ______ hybridized. [1]

In the following questions a statement of Assertion (A) followed by Reason (R) is given. Use the following key to select correct answer :

(a) Both Assertion and Reason are correct but Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

(c) Both Assertion and Reason are incorrect.

(d) Assertion is not correct but Reason is correct.

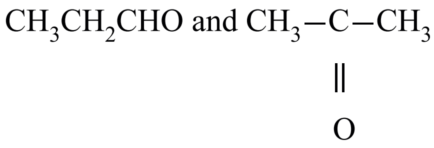
4. Assertion : Carbocations are planar in nature. [1]

   Reason : Carbocations are sp^2 Hybridized.

5. Assertion : All the carbon atoms of But-2-ene lie in are plane.

   Reason : All the carbon atoms in But-2-ene are sp^2 hybridized.

6. (i) What type of isomerism is exhibited by the following pair of compounds? [1]



$$CH_3CH_2CHO \text{ and } CH_3-\underset{\displaystyle O}{\underset{\|}{C}}-CH_3$$

   (ii) Give one example each of nucleophile and electrophile.

7. (i) Arrange the following in increasing order of stability:

$(CH_3)_3C^-$, $CH_3-\overset{-}{C}H-CH_3$, $CH_3CH_2^-$

(ii) Differentiate between inductive and electromeric effect.

8. (i) When do we use hot water funnel for filtration?

(ii) How will you separate a mixture of two organic compounds which have different solubilities in the same solvent?

(iii) An organic liquid decomposes below its boiling point. How will you purify it?

9. Draw the resonating structures of (a) Phenol (b) Benzaldehyde.

10. Arrange the following in the order of property indicated against each set:

(i) $-C_6H_5$, $-NO_2$, $-COOH$, $-I$, $-F$, $-CH_3$, $-C_2H_5$

(In the increasing order of –I effect)

(ii) $CH_3CH_2CH_2CH_2^+$, $(CH_3)_3C^+$, $CH_3CH_2CH_2\overset{+}{C}HCH_3$

(In the order of increasing stability)

(iii) $-COOH$, $-CONH_2$, $-CHO$

(In the increasing priority order if present in same molecule)

(iv) $HCOOH$, CH_3COOH, $ClCH_2COOH$

(Increasing order of acidic strength)

(v) $O_2NCH_2CH_2O^-$, $CH_3CH_2O^-$

(species having greater stability)

*****

UNIT TEST-II

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. Which of the following can act as an electropliile? [1]

   (a) CN^- (b) OH^- (c) H_2O (d) BF_3

2. The most stable free radical among the following is : [1]

   (a) $C_6H_5\,CH_2\,\overset{\circ}{C}H_2$ (b) $C_6H_5\,\overset{\circ}{C}H\,CH_3$

   (c) $CH_3\,\overset{\circ}{C}H_2$ (d) $CH_3\,\overset{\circ}{C}H\,CH_3$

3. What is the other name for no band resonance? [1]

In the following questions a statement of Assertion (A) followed by Reason (R) is given. Use the following key to select correct answer :

(a) Both Assertion and Reason are correct but Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

(c) Both Assertion and Reason are incorrect.

(d) Assertion is not correct but Reason is correct.

4. Assertion : In $CH_2 = C = CH_2$ all the carbon atoms are sp^2 hybridized. [1]

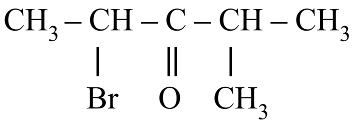
   Reason : All the hydrogen atoms lie in one plane.

5. Assertion : Glycerol is purified by distillation under reduce pressure.

   Reason : Organic compound in liquid state are purified by distillation.

6. Give IUPAC name of the following compounds : [2]

(a) $C_6H_5\,CH_2\,CH_2OH$

(b) 

$CH_3-\underset{\substack{|\\ Br}}{CH}-\underset{\substack{||\\ O}}{C}-\underset{\substack{|\\ CH_3}}{CH}-CH_3$

7. Discuss the principle of steam distillation. [2]

8. Explain the reason for the fusion of an organic compound with metallic sodium for testing nitrogen sulphur and halogens. [2]

9. What are the various conditions essential for resonance? [3]

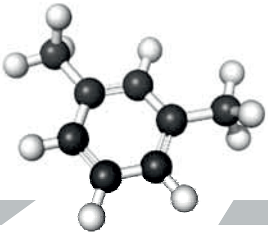
10. Suggest a method to separate the constituents from the following mixture : [5]

(a) Mixture of two miscible liquids.

(b) A mixture of oil and water.

(c) A mixture of plant pigments.

*****



Chapter - 9

Hydrocarbons

FAST TRACK : QUICK REVISION

Hydrocarbons are the organic compounds containing carbon and hydrogen only. Depending upon the types of carbon-carbon bonds present, hydrocarbons can be classified into three categories- (i) Saturated (ii) Unsaturated (iii) Aromatic hydrocarbons.

Saturated hydrocarbons contain carbon-carbon muiltiple bonds–double bonds, triple bonds or both.

ALKANES : Saturated open chain hydrocarbons containing carbon-carbon single bonds. These are inert under normal conditions i.e. do not react with acids, bases and other reagents. Alkanes exhibit Chain isomerism, Position isomerism and conformational isomerism.

General methods of preparation of alkanes :

1. **From Unsaturated hydrocarbons :** By hydrogenation in the presence of platinium, palladium or nickel as catalyst.

General Chemical Equation :

$$R-CH=CH_2+H_2 \xrightarrow[\Delta]{Ni} R-CH_2-CH_3$$

[Where R is H or Alkyl group]

e.g.

$$\underset{\text{Ethene}}{CH_2=CH_2}+H_2 \xrightarrow[\Delta]{Ni} \underset{\text{Ethane}}{CH_3-CH_3}$$

$$\underset{\text{Propyne}}{CH_3-C\equiv CH}+H_2 \xrightarrow[\Delta]{Ni} \underset{\text{Propene}}{CH_3-CH=CH_2}+H_2 \xrightarrow[\Delta]{Ni} \underset{\text{Propane}}{CH_3-CH_2-CH_3}$$

2. **From alkyl halides :** on reduction with Zinc and dilute hydrochloric acid

$$\underset{\text{Chloromethane}}{CH_3-Cl+H_2} \xrightarrow{Zn,\,H^+} \underset{\text{Methane}}{CH_4}+HCl$$

$$\underset{\text{Chloroethane}}{C_2H_5-Cl+H_2} \xrightarrow{Zn,\,H^+} \underset{\text{Ethane}}{C_2H_6}+HCl$$

$$\underset{\text{1-Chloropropane}}{CH_3CH_2CH_2Cl} + H_2 \xrightarrow{Zn,\ H^+} \underset{\text{Propane}}{CH_3CH_2CH_3} + HCl$$

3. **From alkyl halides by Wurtz reaction :** Reaction of alkyl halide with sodium in dry ether, useful only for the preparation of symmetrical alkanes.

$$\underset{\text{Bromoethane}}{CH_3CH_2Br} + 2Na + Br\,CH_2CH_3 \xrightarrow{\text{Dry Ether}} \underset{\text{Butane}}{CH_3CH_2CH_2CH_3}$$

$$\underset{\text{1-Bromopropane}}{CH_3CH_2CH_2Br} + 2Na + BrCH_2CH_2CH_3 \xrightarrow{\text{Dry Ether}} \underset{\text{n–Hexane}}{CH_3CH_2CH_2CH_2CH_2CH_3}$$

$$\underset{\text{1-Bromopropane}}{CH_3CH_2\ Br} + 2Na + \underset{\text{1–Bromoethane}}{Br\ CH_2CH_3} \xrightarrow{\text{Dry Ether}} \underset{\text{n–Pentane (desired)}}{CH_3CH_2CH_2CH_3}$$

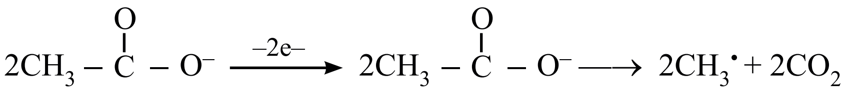
4. **From Carboxylic acids :** By decarboxylation with soda lime

$$\underset{\text{Sodium acetate}}{CH_3COONa} + \underset{\text{Sodium Hydroxide}}{NaOH} \xrightarrow[CaO]{\Delta} \underset{\text{Methane}}{CH_4} + \underset{\text{Sodium}}{Na_2CO_3}$$

5. **By Kolbe's electrolytic method :** Electrolysis of an aqueous solution sodium or potassium salt of carboxylic acid. Alkane containing even number of carbon atoms is formed at anode.

$$2CH_3COO\ Na + 2H_2O \xrightarrow{\text{Electrolysis}} CH_3 - CH_3 + 2CO_2 + H_2 + 2NaOH$$

At Anode : (Oxidation)



$$2CH_3 - \overset{\overset{\large O}{|}}{C} - O^- \xrightarrow{-2e-} 2CH_3 - \overset{\overset{\large O}{|}}{C} - O^- \longrightarrow 2CH_3^{\bullet} + 2CO_2$$

$$2CH_3^{\bullet} \longrightarrow CH_3 - CH_3$$

At Cathode : (Reduction)

$$2H_2O \longrightarrow 2OH^- + 2H^+$$

$$2H^+ + 2e^- \longrightarrow H_2$$

Physical Properties of alkanes :

1. Boiling point of alkanes decreases on branching due to decrease in surface area of molecule with branching which decreases magnitude of van der Waal's forces of attraction.
2. Alkanes being non-polar in nature are soluble in non-polar solvents.

Chemical properties of Alkanes :

- Alkanes undergo substitution reactions.
  e.g., Halogenation, Nitration, Sulphonation.

  Halogenation : For example Chlorination of methane

$$CH_4 + Cl_2 \longrightarrow \underset{\text{Chloromethane}}{CH_3Cl} + HCl$$

$$CH_3Cl + Cl_2 \longrightarrow \underset{\text{Dichloromethane}}{CH_2Cl_2} + HCl$$

$$CH_2Cl_2 + Cl_2 \longrightarrow \underset{\text{Trichloromethane}}{CHCl_3} + HCl$$

$$CHCl_3 + Cl_2 \longrightarrow \underset{\text{Tetrachloromethane}}{CCl_4} + HCl$$

Rate of reaction of alkanes with halogens is $F_2 > Cl_2 > Br_2 > I_2$. Rate of replacement of hydrogen in alkanes is 3° > 2° > 1°. Fluorination is too violent to be controlled. Iodination is reversible and it is therefore carried out in the presence of oxidising agent like HNO_3.

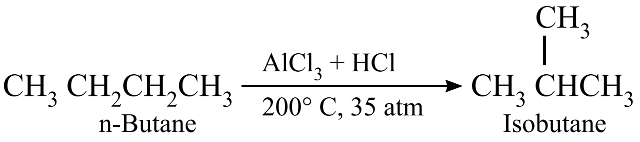
Mechanism of halogenation : Free radical mechanism

$Cl - Cl \xrightarrow{h\nu} Cl\bullet + \bullet Cl$	Initiation
$Cl\bullet + CH_4 \longrightarrow H - Cl + \bullet CH_3$	Propagation
$\bullet CH_3 + Cl - Cl \longrightarrow CH_3\,Cl + \bullet Cl$	Propagation
$Cl\bullet + \bullet CH_3 \longrightarrow CH_3\,Cl$	Termination

Combustion : Complete combustion gives carbon dioxide and water.

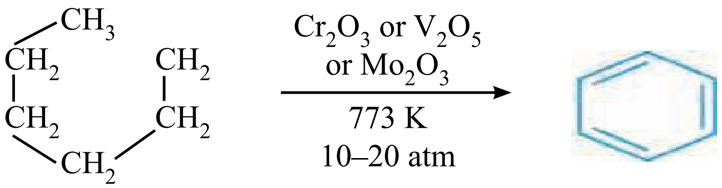
$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Isomerisation :



$$\underset{\text{n-Butane}}{CH_3\,CH_2CH_2CH_3} \xrightarrow[200^\circ\,C,\ 35\,atm]{AlCl_3 + HCl} \underset{\text{Isobutane}}{CH_3\,\overset{\overset{CH_3}{|}}{C}HCH_3}$$

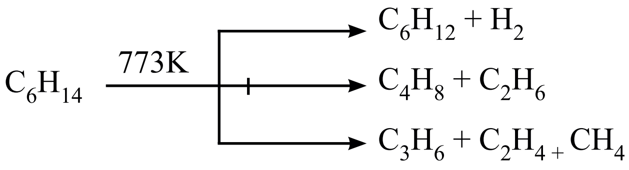
Aromatisation :



(vi) **Reaction with steam**

$$CH_4 + H_2O \xrightarrow[\Delta]{Ni} CO + 3H_2$$

Pyrolysis : Decomposition of higher alkanes to lower alkanes on heating.



$$C_6H_{14} \xrightarrow{773K} \begin{cases} C_6H_{12} + H_2 \\ C_4H_8 + C_2H_6 \\ C_3H_6 + C_2H_4 + CH_4 \end{cases}$$

Conformations :

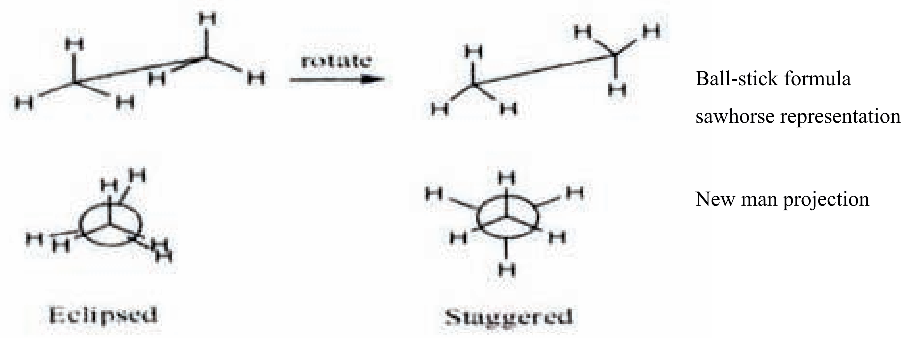
Different spatial arrangement of atoms arising due to rotation around C-C single bond.

Conformation of ethane, CH_3CH_3

Two conformational isomers or conformers.

Eclipsed form = all hydrogen atoms nearest to each other.

Staggered form = all hydrogen atoms are farthest apart.



Stability of eclipsed conformation is least while staggered conformation is most stable. The energy difference between two extreme forms is $12.5 kJmol^{-1}$. Due to this small energy difference the two forms are easily inter-convertible at ordinary temperature and cannot be separated and isolated.

ALKENES

These are unsaturated non-cyclic hydrocarbons which have.sp^2 -hybridisation with 120° bond angle.

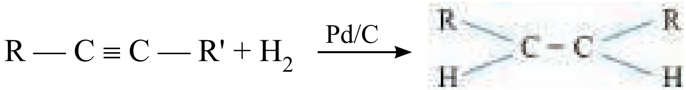
Alkenes are also called olefins [oil.forming] which indicates their high reactive nature.

Alkenes have general formula $C_n H_{2n}$, where n = 2, 3, 4, ...

C_2H_4 (ethene), C_3H_6 (propene), etc.

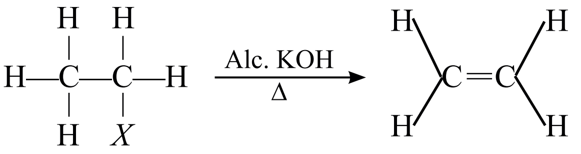
Methods of Preparation of Alkenes

(i) From alkynes



cis-alkene

(ii) From alkyl halide [Dehydrohalogenational]

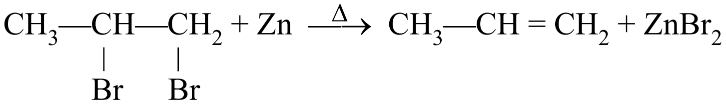


[X = Cl, Br, I]

[β-Elimination product]

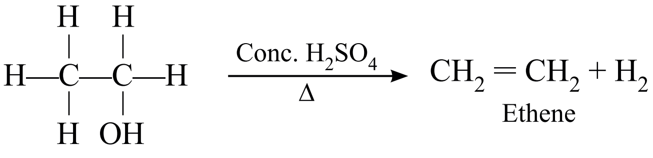
(iii) From *vicinal* dihalides

$$CH_2Br—CH_2Br + Zn \xrightarrow{\Delta} CH_2 = CH_2 + ZnBr_2$$



$$\underset{\quad\ \ Br\quad\ \ Br}{CH_3—CH—CH_2} + Zn \xrightarrow{\Delta} CH_3—CH = CH_2 + ZnBr_2$$

(iv) From alcohols by acidic dehydrogenation



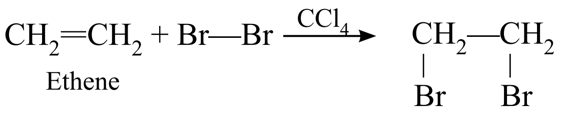
$$H—\overset{H}{\underset{H}{C}}—\overset{H}{\underset{OH}{C}}—H \xrightarrow[\Delta]{Conc.\ H_2SO_4} CH_2 = CH_2 + H_2$$

Ethanol

Ethene

Chemical Properties of Alkenes :

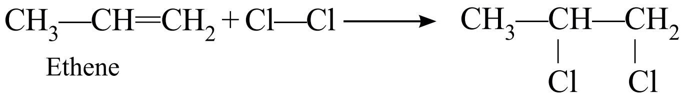
1. Addition of Halogens :



$$CH_2{=}CH_2 + Br{-}Br \xrightarrow{CCl_4} \underset{Br}{\underset{|}{CH_2}}{-}\underset{Br}{\underset{|}{CH_2}}$$

Ethene

1, 2-Dibromoethane



$$CH_3{-}CH{=}CH_2 + Cl{-}Cl \longrightarrow CH_3{-}\underset{Cl}{\underset{|}{CH}}{-}\underset{Cl}{\underset{|}{CH_2}}$$

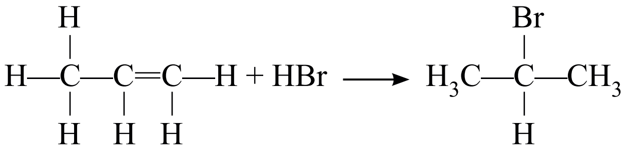
Ethene

1, 2-Dichloropropane

2. Addition of hydrogen halides HCl, HBr, HI :

Addition of hydrogen halides HCl, HBr, HI : Add up to alkenes to form alkyl halides as per their reactivity order in HI>HBr>HCl.

Addition reaction of HBr to unsymmetrical alkenes (Markownikov's rule) According to Markownikovs' rule, the negative part of the addendum (adding molecule) gets attached to that carbon atom which possesses lesser number of hydrogen atoms.



$$H{-}\overset{H}{\overset{|}{\underset{H}{\underset{|}{C}}}}{-}\underset{H}{\underset{|}{C}}{=}\underset{H}{\underset{|}{C}}{-}H + HBr \longrightarrow H_3C{-}\overset{Br}{\overset{|}{\underset{H}{\underset{|}{C}}}}{-}CH_3$$

2-Bromopropane

Anti Markownikov addition or peroxide effect or Kharasch effect in the presence of organic peroxide, addition of only HBr molecule on unsymmetrical alkene takes place contrary to the Markownikov's rule. Peroxide effect is not observed in case of HF, HCl and HI.

$$CH_3{-}CH{=}CH_2 + HBr \xrightarrow[(C_6H_5CO)_2O_2]{} CH_3CH_2CH_2Br$$

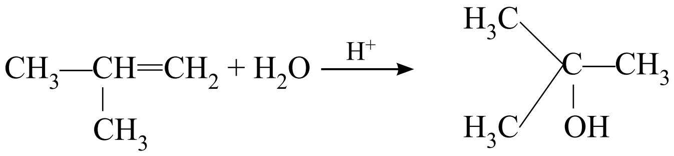
1-bromopropane

3. Addition of sulphuric acid

$$CH_3{-}CH{=}CH_2 + \underset{\text{[cold and conc.]}}{HOSO_2OH} \rightarrow CH_3CH(OSO_3H)\,CH_3$$

propyl hydrogen sulphate

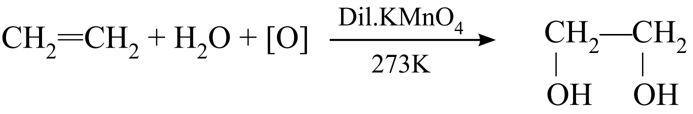
4. Addition of water



2-Methylpropene

2-Methylpropan-2-ol

5. **Oxidation :** Alkenes decolourise cold dilute aqueous solution potassium permanganate (Baeyer's reagent). It is used as a test for unsaturation.



$$CH_2{=}CH_2 + H_2O + [O] \xrightarrow[273K]{Dil.KMnO_4} \underset{OH}{CH_2}{-}\underset{OH}{CH_2}$$

Ethane–1,2-diol
or glycol

Acidic $KMnO_4$ or acidic $K_2Cr_2O_7$ oxidises alkenes to ketones and/or acids depending upon the nature of alkene and the experimental conditions.

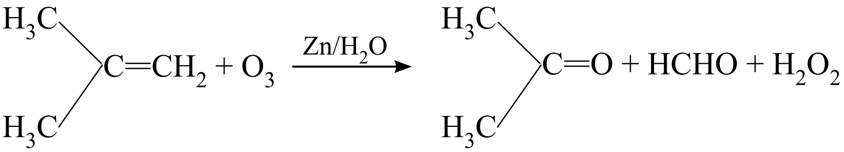
$$(CH_3)_2C{=}CH_2 \xrightarrow{KMnO_4/H^+} (CH_3)_2CO + CO_2 + H_2O$$

2-Methyl propene Propane-2-one

$$CH_3{-}CH{=}CH{-}CH_3 \xrightarrow{KMnO_4/H^+} 2CH_3COOH$$

But-2-ene Ethanoic acid

6. **Ozonolysis :** Reaction of ozone with alkene to form ozonide which on subsequent reductive cleavage with zinc dust and water give carbonyl compounds (aldehydes & ketones).

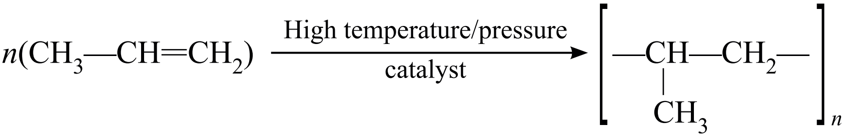


$$Zn + H_2O_2 \longrightarrow ZnO + H_2O$$

7. **Polymerization**

$$n(CH_2{=}CH_2) \xrightarrow[\text{catalyst}]{\text{High temperature/pressure}} {-}[CH_2{-}CH_2{-}]_n$$

polythene



$$n(CH_3{-}CH{=}CH_2) \xrightarrow[\text{catalyst}]{\text{High temperature/pressure}} \left[{-}\underset{CH_3}{CH}{-}CH_2{-}\right]_n$$

ALKYNES

These are unsaturated hydrocarbons with general formula C_nH_{2n-2} e.g., C_2H_2 (ethyne), C_3H_4 (propyne).

Alkynes also exhibit electrophilic addition reaction but less reactive than alkenes because the dissociation of π-electron cloud requires more energy.

H–C≡C–H contins 3σ and 2π -bonds and bond length is 120 pm. In acetylene. H–C–C bond angle is 180°.

Methods of Preparation of Alkynes

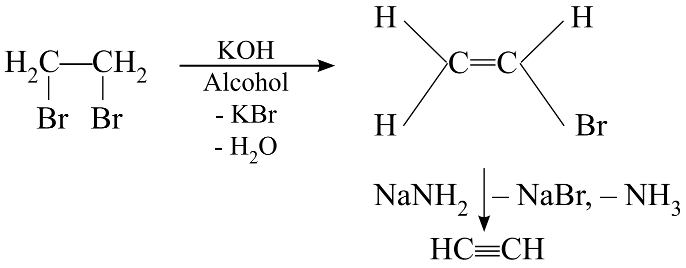
1. From calcium carbide

$$CaCO_3 \longrightarrow CaO + CO_2$$

$$CaO + C \longrightarrow CaC_2 + CO$$

$$CaC_2 + 2H_2O \longrightarrow Ca(OH)_2 + C_2H_2$$

2. From vicinal dihalides



3. From tetrahalides

$$Br_2CH{-}CHBr_2 + 2Zn \xrightarrow[\Delta]{CH_3OH} H{-}C{\equiv}C{-}H + 2ZnBr_2$$

Physical Properties of Alkynes :

1. The first two members are gases next eight members ($C_5 - C_{12}$) are liquids and higher members are solids.
2. They are all colourless and odourless with the exception of acetylene which has slightly garlic odour due to the presence of PH_3 and H_2S as impurities.
3. Alkynes are insoluble in water but soluble in organic solvents like ethers, carbon tetrachloride and benzene.
4. Melting point, boiling point and density increase with increase in molar mass.

Chemical properties of Alkynes

Alkynes show electrophilic as well as nucleophilic addition reactions.

(i) Acidic character of alkyne

$$HC \equiv CH + Na \longrightarrow \underset{\text{monosodium acetylides}}{HC \equiv \bar{C}Na} + \frac{1}{2}H_2$$

$$H - C \equiv CNa^+ + Na \longrightarrow \underset{\text{disodium acetylide}}{Na^+ C \equiv \bar{C}Na^+} + \frac{1}{2}H_2$$

$$CH_3 - C \equiv C - H + NaNH_2 \longrightarrow \underset{\text{sodium propynide}}{CH_3 - C \equiv \bar{C}Na^+} + NH_3$$

These reactions are not shown by alkenes, alkanes and non-terminal alkynes, hence used for distinction between alkane, alkene and alkyne.

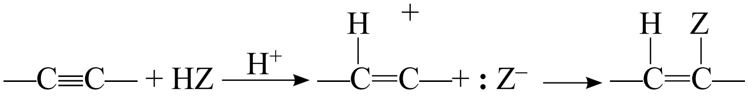
Acetylenic hydrogens are acidic in nature due to 50% s-character in sp-hybridised orbitals. Acidity of alkynes is lesser than water.

Acidic behaviour order

(i) $\underset{sp}{HC{\equiv}CH} > \underset{sp^2}{CH_2 = CH_2} > \underset{sp^3}{CH_3 - CH_3}$

(ii) $HC{\equiv}CH > CH_3{-}C{\equiv}CH > CH_3{-}C{\equiv}C{-}CH_3$

(ii) Electrophilic addition reactions



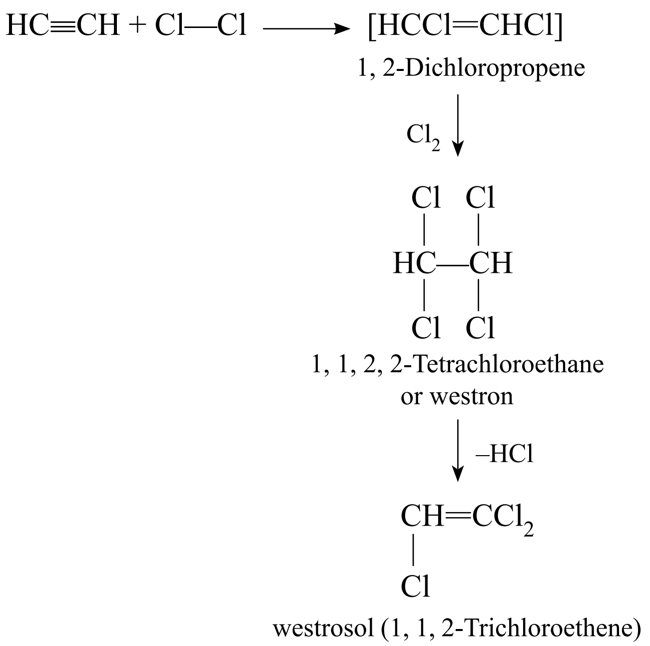
$$-C\equiv C- + HZ \xrightarrow{H^+} -\overset{\overset{H}{|}}{C}=\overset{+}{C}- + :Z^- \longrightarrow -\overset{\overset{H}{|}}{C}=\overset{\overset{Z}{|}}{C}-$$

vinyl cation

The addition product formed depends upon the stability of vinylic cation. Addition on unsymmetrical alkynes takes place according to Markovnikov's rule.

$$CH_3{-}C{\equiv}C{-}H + H_2 \xrightarrow{Pt/Pd/Ni} \underset{\text{propene}}{[CH_3{-}CH{=}CH_2]} \xrightarrow{H_2} \underset{\text{propene}}{CH_3CH_2CH_3}$$

Addition of halogens



$$HC\equiv CH + Cl{-}Cl \longrightarrow [HCCl{=}CHCl]$$

1, 2-Dichloropropene

$$\Big\downarrow Cl_2$$

$$\begin{array}{cc} Cl & Cl \\ | & | \\ HC & {-}CH \\ | & | \\ Cl & Cl \end{array}$$

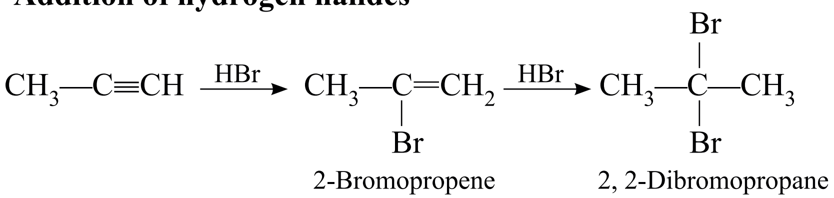
1, 1, 2, 2-Tetrachloroethane
or westron

$$\Big\downarrow -HCl$$

$$\begin{array}{l} CH{=}CCl_2 \\ | \\ Cl \end{array}$$

westrosol (1, 1, 2-Trichloroethene)

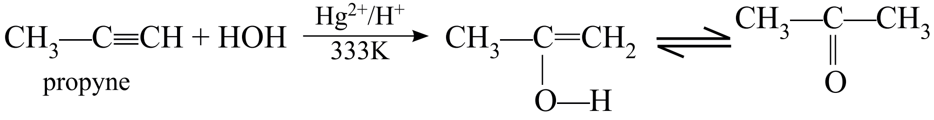
Addition of hydrogen halides



$$CH_3{-}C\equiv CH \xrightarrow{HBr} \underset{\underset{Br}{|}}{CH_3{-}C{=}CH_2} \xrightarrow{HBr} CH_3{-}\overset{\overset{Br}{|}}{\underset{\underset{Br}{|}}{C}}{-}CH_3$$

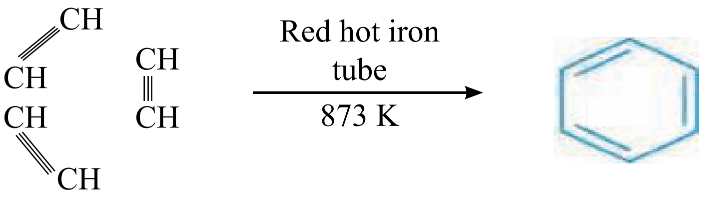
2-Bromopropene 2, 2-Dibromopropane

Addition for water



$$\underset{\text{propyne}}{CH_3{-}C\equiv CH} + HOH \xrightarrow[333K]{Hg^{2+}/H^+} CH_3{-}\underset{\underset{O-H}{|}}{C}{=}CH_2 \rightleftharpoons CH_3{-}\underset{\underset{O}{\|}}{C}{-}CH_3$$

(iii) Cyclic polymerisation of ethyne

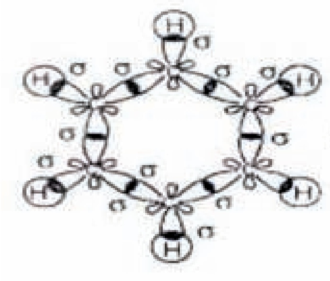


AROMATIC HYDROCARBONS

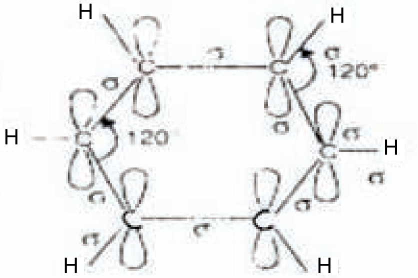
These hydrocarbon are also known as arenes. The parent member of the family aromatic hydrocarbons is benzene.

Aromatic compounds containing benzene ring are known as benzenoids.

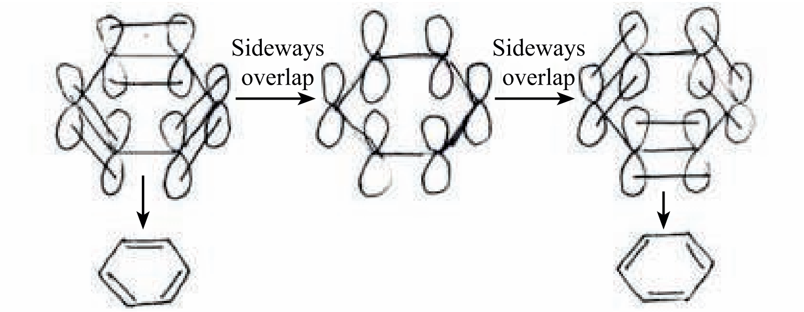
Structure of benzene : Hexagonal ring of carbon atoms with alternate single and double bonds. Each carbon atom is sp^2 hybridised. Planar ring, bond angle 120°. All C-C bond lengths are equal due to complete delocalisation of π electrons.



Formation of C-C and C-H sigma bonds

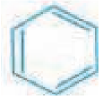


Sigma skeleton of benzene molecule



HUCKEL'S RULE

- **Huckel's rule, (based on calculations) :** a planar cyclic molecule with alternating double and single bonds has aromatic stability if it has $(4n+2)\pi$ electrons (n is 0, 1, 2, 3, 4)
- For $n = 1 : 4n+2 = 6$; benzene is stable and electrons are delocalized.



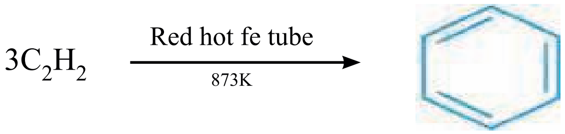
Benzene

Three double bonds;

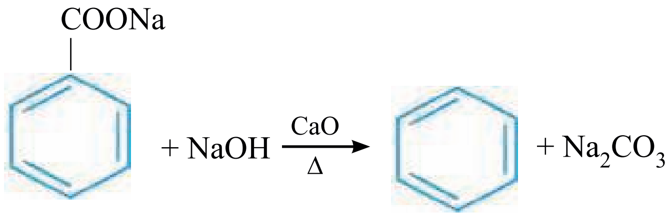
six π electrons

• METHODS OF PREPARATION

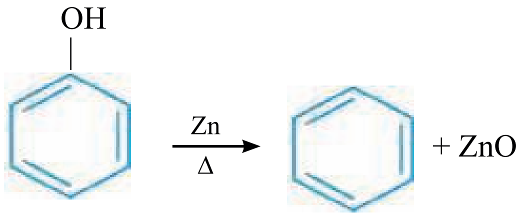
1. Cyclic polymerisation of ethyne



2. Decarboxylation of aromatic acids



3. Reduction of phenol



Physical Properties of Benzene :

(i) Aromatic hydrocarbons are non-polar molecules and are usually colourless liquids or solids with a characteristic aroma.

(ii) Aromatic hydrocarbons are immiscible with water but readily miscible with organic solvents.

(iii) Aromatic compounds burn with sooty flame.

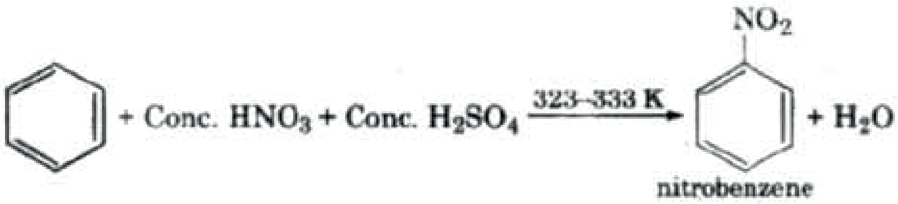
Chemical Reactions of Benzene :

(i) Benzene gives electrophile substitution reactions.

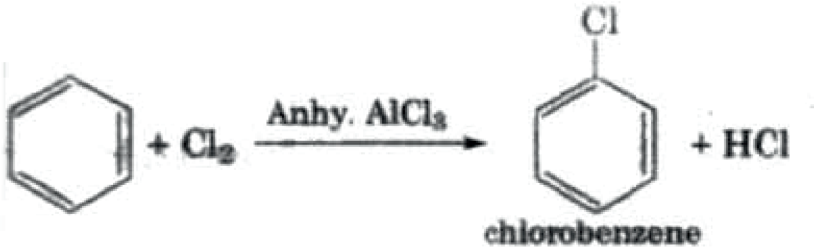
(ii) According to experimental evidences, electrophile substitution reaction involve following three steps :

- Generation of electrophilie
- Formation of carbocation intermediate.
- Removal of proton from the carbocation intermediate.

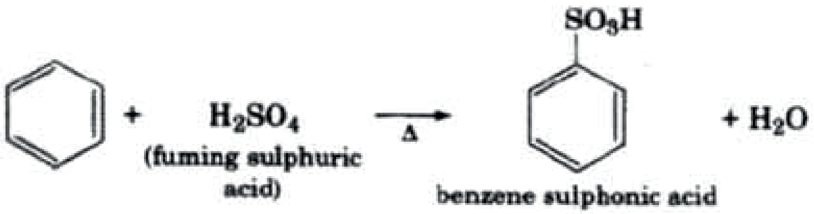
(i) Nitration



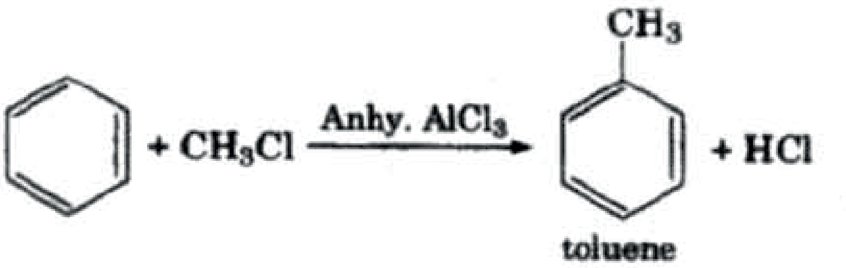
(ii) Halogenation



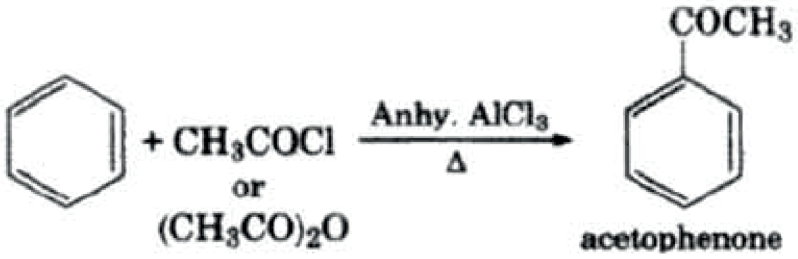
(iii) Sulphonation



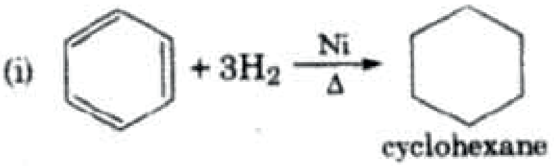
(iv) Friedel-Craft's alkylation reaction

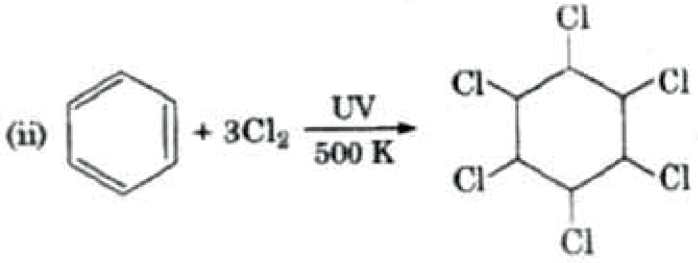


(v) Friedel-Crafts acylation reaction



Benzene also undergoes addition reactions e.g.





benzene bexachloride or 666
(BHC or Gammexane or lindane)

- **COMBUSTION**

$2C_6H_6 + 15O_2 \longrightarrow 12CO_2 + 6H_2O$

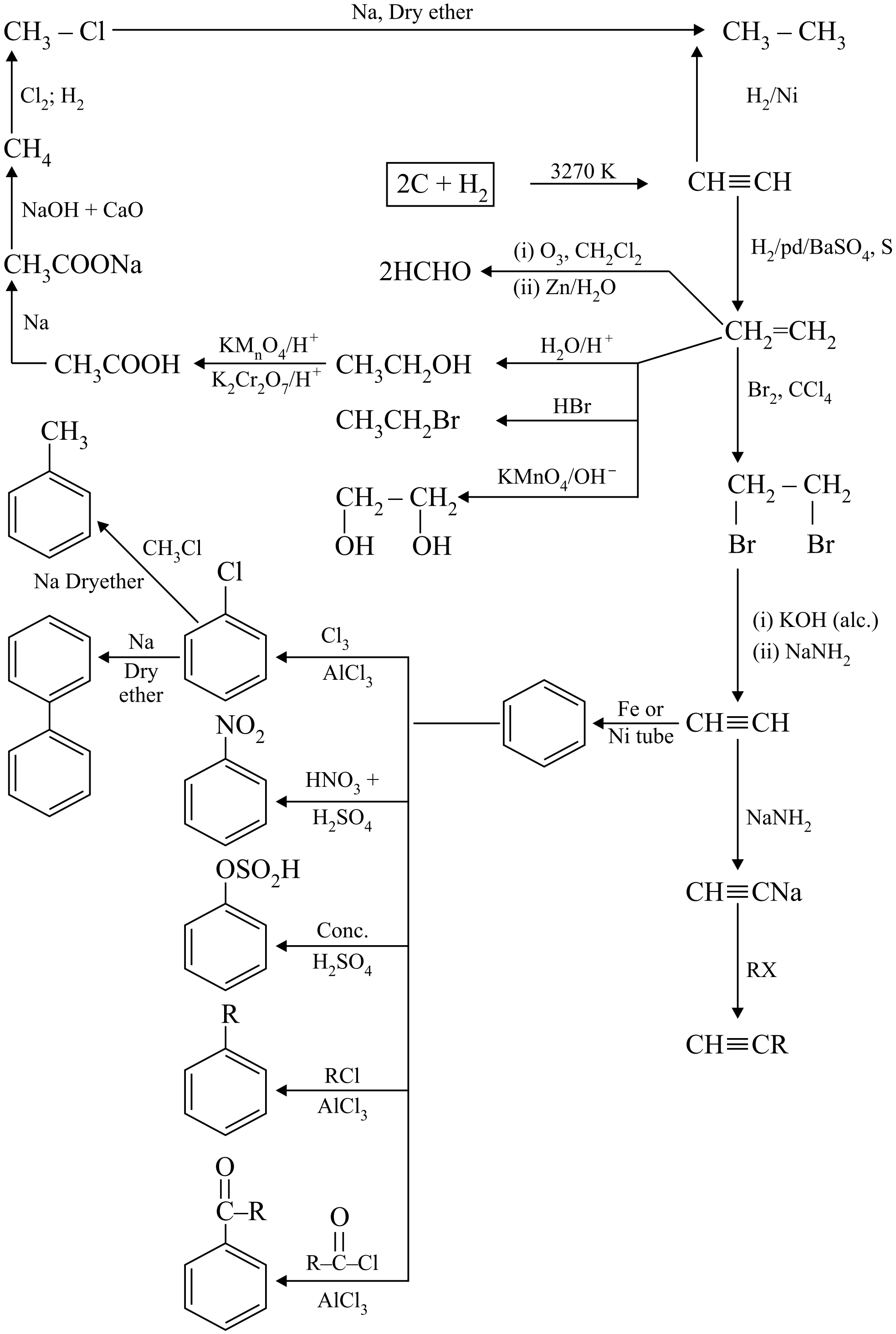
Directing influence of substituents in monosubstituted benzene

(i) **Ortho and para directing groups :** Ring activating groups *e.g.,* $NH_2, -CH_3, -C_2H_5, -OCH_3$ etc. (+ R effect)

(ii) **Meta directing groups :** Ring deactivating groups

e.g. - $NO_2, -CN, -CHO, -COOH, -SO_3H$ (– R effect).

MIND MAP : HYDROCARBON



CASE BASED STUDY - QUESTIONS

PASSAGE–I

Presence of double bond in an alkene is tested by reacting with either water or with dilute solution of Bayer's regent. The position of double bond is located with the help of ozonalysis. It is done by joining together the product of ozonalysis which are carbonyl compounds at their carbonyl carbon atoms by double bond. One mole of ozone is used in the ozonalysis reaction per mole of double bond in a particular alkene.

1. An alkene upon ozonalysis will give CH_3CH_2 CHO and CH_3COCH_3. Give IUPAC name and structure of the given alkene.

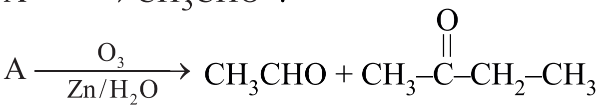
2. Complete the given chemical equation.

   $$CH_3 - CH = CH - CH_3 + O_3 \xrightarrow[H_2O]{Zn} A$$

3. On ozonolysis one molecule of a hydrocarbon produces two molecules of ethanal and one molecule of ethanedial. Give the structure on ozonolysis of hydrocarbon.

4. An alkene give 5-keto-2-methyl hexanal. Give IUPAC name of the given alkene.

5. (a) $A \xrightarrow{O_3} CH_3CHO$?

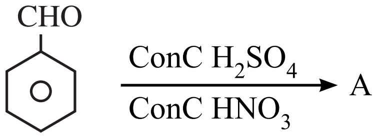
   (b) 

   $$A \xrightarrow[Zn/H_2O]{O_3} CH_3CHO + CH_3-\overset{\overset{\large O}{||}}{C}-CH_2-CH_3$$

PASSAGE–II

Benzene responds to electrophilic substitution reactions. However some reagents are needed in these reactions in order to generate the attaching electrohiles. The directive influence in disubstitution taking place in the ring depends upon the nature of the group already present. An activating group in creases the electron density at the ortho and para positions. The deactivating group decreases the same at these positions thereby healing the metal position comparatively a point of high electron density.

1. Complete the following chemical equation

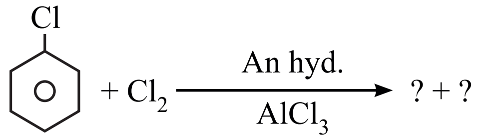


2. Explain why ($-NO_2$) group is m-directing & reactivating group.

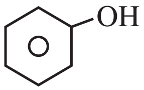
3. Arrange the following compound in increase order of reactivity towards electrophilic substitution reaction



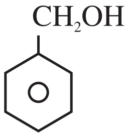

4. Write the product in the following reaction

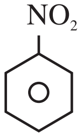


5. Which of the following is most reactive to electrophilic attack.


(a)


(b)


(c)


(d)

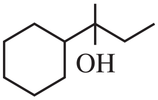
MULTIPLE CHOICE QUESTIONS (MCQ)

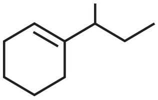
1. Which of the following has zero dipole moment?
   - (a.) cis-But-2-ene
   - (b) trans-But-2-ene
   - (c) But-1-ene
   - (d) 2-Methylprop-1-ene

2. Bond length of (I) ethane, (II) ethene, (III) Acetylene, (IV) Benzene follows the order:
   - (a) I > II > III > IV
   - (b) I > II > IV > III
   - (c) I > IV > II > III
   - (d) III > IV > II > I

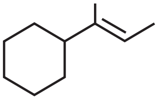
3. The methyl group in benzene ring is:
   - (a) Ortho directing
   - (b) Ortho and meta directing
   - (c) Para directing
   - (d) Ortho and para directing

4. Which of th following is not the product of dehydration of

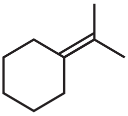




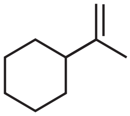
(a)



(b)



(c)



(d)

5. In the hydrocarbon :

$$\overset{6}{H_3C} - \overset{5}{CH} = \overset{4}{CH} - \overset{3}{CH_2} - \overset{2}{C} \equiv \overset{1}{CH}$$

The state of hybridization of carbon 1, 3 and 5 are in the order

(a) sp, sp^2, sp^3 (b) sp^3, sp^2, sp

(c) sp^2, sp, sp^3 (d) sp, sp^3, sp^2

6. Action of acetylene on dilute H_2SO_4 gives:

(a) Acetic acid (b) Acetone

(c) Acetaldehyde (d) Ethyl alcohol

7. Which of the following compounds will exhibit cis-trans (geometrical) isomerism?

(a) Butanol (b) 2-Butyne

(c) 2-Butenol (d) 2-Butene

8. Basic strength of :

$H_3C\ CH$, $H_2C = CH$ and $HC \equiv C$ is in the order of

(a) I > III > II (b) I > II > III

(c) II > I > III (d) III > II > I

9. Reaction of hydrogen bromide with propene in the absence of a peroxide is a/an

(a) free radical reaction (b) nucleophilic substitution

(c) electrophilic addition (d) nucleophilic substitution

10. Among the following compounds, the one which is most reactive towards electrophilic nitration is:

(a) Benzoic acid (b) Nitrobenzene

(c) $\ominus$ Toluene $\ominus$ $\ominus$ (d) Benzene

ANSWERS: 1. b 2. c 3. d 4. a 5. d 6. c 7. d 8. c 9. c 10. c

FILL IN THE BLANKS

1. The addition of HBr to an unsymmetrical alkene takes place in accordance with ___________ rule.
2. Benzene favours ___________ substitution reaction.
3. The Dipole moment of Benzene is ___________.
4. The nitro group in the benzene nucleus is ________ directing. It __________ the reactivity of the benzene ring.
5. Melting point and boiling point increase as the molar masses _______.

ANSWERS : 1. Markownikov's 2. electrophilic 3. Zero
4. Meta, decreases 5. Increase

TRUE AND FALSE TYPE QUESTIONS

1. Alkanes mainly undergo substitution reactions using the free-radical mechanism.
2. The decreasing order of boiling points among the isomeric pentanes is neo > iso > n.
3. The acidic character of three types of hydrocarbons follows the order alkanes > alkenes > alkynes.
4. The peroxide effect is observed only in addition of HBr, and not with HCl and HI.
5. Wurtz reaction is suitable for the preparation of both symmetrical and unsymmetrical alkanes.
6. For a compound to be aromatic it must have $(4n + 2)\pi$ electrons.
7. Benzene has planar structure.
8. The benzene molecule has two different carbon-carbon bond lengths, corresponding to alternate single and double bonds.
9. In Friedel-Crafts reaction, $AlCl_3$ is an electrophile.
10. An electron-donating substituent in benzene ring gives a meta product.

ANSWERS: 1. T 2. F 3. F 4. T 5. F 6. T 7.T 8. F 9. F 10. F

MATCH THE COLUMNS

Match the statements (a,b,c,d) in column I with the statements (i, ii, iii, iv) in column II.

1.	**Column I**	**Column II**	**Column III**
	a. CH_4	i. sp^2	e. Ozonalysis
	b. $CH_2 = CH_2$	ii. sp^3	f. Oxidising agent
	c. $CH \equiv CH$	iii. sp^3d	g. Saturated nature
	d. PCl_5	iv. sp	h. Un Saturated nature

2.

Column I	Column II
a. Alkanes	i. Saturated nature
b. Alkenes	ii. Ozonolysis
c. Alkynes	iii. Geometrical isomerism
d. Arenes	iv. Aromatic character

ANSWERS: 1. a. → ii. - g. b. → i.-h.-e. c. → i.-h.-e. d. → iii.-f.
2. a. → i. b. → ii., iii. c. → ii. d. → i., iv.

ASSERTION-REASON TYPE QUESTIONS

Type 1. The questions given below consist of Assertion(A) and Reason (R). Use the following key to select correct answer.

(a) If both assertion and reason are correct and reason is correct explanation for assertion.

(b) If both assertion and reason are correct but reason is not correct explanation for assertion.

(c) If assertion is correct but a reason is incorrect.

(d) If assertion and reason both are incorrect.

1. **Assertion:** The IUPAC name of $CH_3CH = CH - C \equiv CH$ is pent-3-en-1-yne and not pent-2-en-4-yne.

   Reason: While deciding the locants of double and triple Bonds, lowest sum rule is always followed.

2. **Assertion:** Tropylium cation  is aromatic in character.

   Reason: The only property which decides the aromatic character is its planar nature.

3. **Assertion:** Friedel-craft reaction between benzene and acetic anhydride in the presence of anhydrous $AlCl_3$ yields acetophenone and not poly-substituted products.

   Reason: Acetophenone formed poisons the catalyst preventing further reaction.

4. **Assertion:** But-1-ene on reacting with HBr in the presence of peroxide, products 1-bromobutane.

   Reason: It involves the formation of a primary free radical.

5. **Assertion:** Cyclopentadienyl anion is aromatic in nature.

   Reason: Cyclopentaclienyl anion has six π-electrons.

6. **Assertion:** Benzene reacts with chlorine in the form of light to form BHC.

   Reason: BHC is also called gammexane or 666.

7. **Assertion:** All the hydrogen atoms in $CH_2 = C = CH_2$ lie in one plane.
   Reason: All the carbon atoms in it are sp^2 hybridised.
8. **Assertion:** Propene reacts with HBr in the presence of benzoyl peroxide to yield 2-bromopropane.
   Reason: In the presence of peroxide, the addition of HBr to propene follows ionic mechanism.
9. **Assertion:** Benzene does not decolourise bromine water.
   Reason: Benzene is stabilised by resonance due to delocalisation of π electrons.
10. **Assertion:** Acidity of C-H bond decreases in the order:
    $HC \equiv CH > H_2C = CH_2 > H_3C - CH_3$
    Reason: Greater the percentage s-character, more is the acidity of C—H bond.

ANSWERS: 1.a 2.c 3.c 4.c 5.a 6.b 7.c 8.d 9.a 10.a

ONE WORD TYPE QUESTIONS

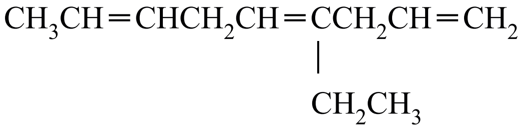
1. What is the state of hybridisation of Carbon atoms in alkanes?
2. What is the number of bonds in But-3-en-1-yne?
3. Name the product formed when Propyne is treated with aqueous H_2SO_4 in the presence of dil.$HgSO_4$.
4. Name the product formed when Benzene reacts with CH_3Cl in the presence of anhydrous aluminium chloride.
5. —COOH is ortho, para directing or Meta directing group?

1-MARK QUESTIONS

1. Write the reaction involved in Kolbe's electrolytic method to prepare ethane.
2. Define term decarboxylation.
3. Why dry ether and not water is used as a solvent in the preparation of alkane by Wurtz reaction?
4. Sodium salt of which carboxylic acid will be needed for the preparation of propane by decarboxylation method?
5. Complete the following reaction:

   $$CH_3Cl + Na \xrightarrow{\text{dry ether}}$$

6. Amongst the following which one has the maximum boiling point?
n-Pentane, iso-pentane, neo-pentane.

7. Write IUPAC name of  $CH_3CH=CHCH_2CH=\underset{\substack{|\\ CH_2CH_3}}{C}CH_2CH=CH_2$

8. Draw the cis and trans isomers of $CHCl=CHCl$.

9. What happens when 2-Bromobutane is being treated with alc. KOH?

10. Name the reagents used to carry out the following conversions:
$CH_3-CH=CH_2 \rightarrow CH_3CH(OH)CH_2OH$

11. Complete the following reaction :
$CH_3-CH=CH_2 + HBr \xrightarrow{\text{Organic peroxide}}$

12. An alkene A on ozonolysis gives a mixture of ethanol and pentan-3-one. Write IUPAC name of element.

13. When alkyne is treated with bromine water then what will be the colour of the product?

14. Why alkynes do not exhibit geometrical isomerism?

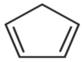
15. Complete the following reaction:
(i) $CH_3C\equiv CH \xrightarrow{H_2O,\ Hg^{2+}/H^+}$?
(ii) $CaC_2 + 2H_2O \longrightarrow$ ____________ + ____________

16. How will you convert ethyne to benzene?

17. Write IUPAC name of $C_6H_5-CH_2-CH_2-CH=CH_2$.

18. Why is benzene extraordinarily stable although it contains three double bonds?

19. Write chemical reaction to exemplify Friedel-Crafts alkylation of benzene.

20. Why  is not aromatic?

21. C-C bond length in benzene is intermediate between C–C and C=C. Why?

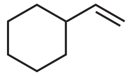
22. Starting from benzene,how would you synthesize m- Bromonitrobenzene.

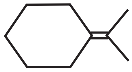
23. Give one example each of o, p-directing group and m-directing group.

24. Complete the reaction:
$C_6H_6 + CH_3COCl \xrightarrow{\text{Anhy. }AlCl_3}$

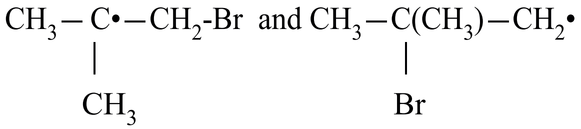
2-MARKS QUESTIONS

1. What effect does branching have on the boiling point of an alkane and why?
2. An alkene A contains three C–C eight C–H and one C–C (π) bond upon ozonlysis A gives two moles of an aldehyde of molar mass 44u. Write the IUPAC name of A.
3. Write the structures of all the alkenes which upon hydrogenation give 2-methyl butane.
4. Wurtz reaction cannot be used for the preparation of unsymmetrical alkanes? Give reason.
5. How can ethene be prepared from (i) Ethanol (ii) Ethyl bromide?
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. Explain with the help of equation : Ozonolysis of propene.
9. What do you understand by peroxide effect (Kharasch effect)?
10. What factor determines the stability of alkene?
11. Out of benzene, *m*-dinitro benzene and toluene, which will undergo nitration most easily and why.
12. Complete the reaction:

(i)  + HBr → ?

(ii)  $+ O_3 \xrightarrow[CH_3COOH]{Zn}$? + ?

13. An alkene on treatment with HBr in presence of peroxide can generate two types of free radicals



$$CH_3-\underset{\underset{CH_3}{|}}{C^\bullet}-CH_2\text{-}Br \text{ and } CH_3-\underset{\underset{Br}{|}}{C(CH_3)}-CH_2^\bullet$$

Predict the final product of the reaction and give reason.

(Hint: Stability of free radicals)

14. What happens when But-2-ene reacts with acidified potassium permanganate solution?
15. You are provided with But-2-yne, how will you convert it into:
    - (i) cis-But-2-ene
    - (ii) trans-But-2-ene

16. An alkene C_4H_8 reacts with HBr both in the presence and absence of peroxide to give the same product. Identify the alkene.

17. Arrange ethane, ethene and ethyne in the order of increasing acidity.

18. Identify A and B in the following reaction:

$$A \xrightarrow{Na} CH{\equiv}CH \xrightarrow[873K]{\text{Red hot iron tube}} B$$

19. Write the structures of the products A and B of the following reactions:

(i) $HC{\equiv}CH \xrightarrow{Na} A \xrightarrow{CH_3Br} B$

(ii) $BrH_2C{-}CH_2Br \xrightarrow{\text{Alc.KOH}} A \xrightarrow{NaNH_2} B$

3-MARKS QUESTIONS

1. Write the structures and name of products obtained in the reaction of sodium with a mixture of 1-Iodo-2-methylpropane and 2-Iodopropane.

2. State Markownikov's rule. Using this rule, write the reaction of propene with (i) HCl & (ii) H_2O.

3. Complete the following reactions:

(i) $CH_3CH_2Br \xrightarrow{\text{Alc.KOH}}$

(ii) $CH_3CH{=}CH_2 + O_3 \xrightarrow{Zn/H_2O}$

(iii) $CH_2{=}CH_2 + H_2O + [O] \xrightarrow{\text{Dil.}KMnO_4}$

4. (i) Write the structure of 3, 4-Dimethylhept-3-ene.

(ii) Name the compounds obtained by ozonolysis of 3-Methylpent-2-ene.

5. Complete the following reactions:

(i) $CH{\equiv}CH \xrightarrow{NaNH_2,CH_3Br} ?$

(ii) $CH{\equiv}CH \xrightarrow{H_2O,HgSO_4/H_2SO_4} ?$

(iii) $CH_3C{\equiv}CH + H_2 \xrightarrow{Pt} ? \xrightarrow{H_2} ?$

6. Write the mechanism of nitration of benzene.

7. Arrange in the order of decreasing relative reactivity with an electrophile and explain:
Toluene, p-Nitrotoluene, 1, 4-Dinitrobenzene

8. What is meant by delocalization of π electrons? How does it affect stability of benzene?

9. What are the conditions for a compound/species to be aromatic according to Huckel's rule?

10. How will you convert benzene into
    (i) Acetophenone
    (ii) m- Chloronitrobenzene?

5-MARKS QUESTIONS

1. Define isomerism. Write all the structural isomers of hexane(C_6H_{14}) and arrange them in increasing order of boiling points.

2. Write short note on (i) Wurtz reaction (ii) Kolbe's electrolysis (iii) Ozonolysis

3. An alkene 'A' of molecular mass '28u' on treatment with bromine gives a product 'B'. The Compound 'B' on further dehalogenation with zinc gives back 'A'. Give the structures of 'A' and 'B' and also the sequence of reactions.

4. An organic compound 'A' with formula C_4H_9Br on treatment with KOH(alc.) gave two isomeric compounds 'B' and 'C' with formula C_4H_8. Ozonolysis of 'B' gave only one product CH_3CHO while 'C' gave two different products. Identify A, B and C.

5. How will you convert Ethyne into (i) 1, 1, 2, 2-Tetrachloroethane (ii) Ethene (iii) Ethanal (iv) Benzene (v) Sodium ethynide

UNIT TEST-I

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. Amongst the following which one has the maximum boiling point and why? n-Pentane , iso-pentane, neo-pentane [1]

2. What is the number of σ and π bonds in But-3-en-1-yne? [1]

3. Action of acetylene on dilute H_2SO_4/dil.$HgSO_4$ gives: [1]

   (a) Acetic acid (b) Acetone (c) Acetaldehyde (d) Ethyl alcohol

In the following questions a statement of Assertion (A) followed by Reason (R) is given. Use the following key to select correct answer :

(a) Both Assertion and Reason are correct but Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

(c) Both Assertion and Reason are incorrect.

(d) Assertion is not correct but Reason is correct.

4. **Assertion :** Benzone reacts with chlorine in the form of light to from BHC.

   Reason : BHC is also called gammexane or 666. [1]

5. **Assertion:** Tropylium cation  is aromatic in character.

   Reason: The only property which decides the aromatic character is its planar nature. [1]

6. Arrange the following alkenes in decreasing order of stability and give reason. [1]

   $CH_3-CH=CH-CH_3$, $CH_2=CH_2$, $CH_3-CH=CH_2$

7. (i) Give a chemical test to distinguish between ethyne and ethene. [2]

(ii) Melting point of cis-But-2-ene is lower than that of trans-But-2-ene. Give reason.

8. Complete the following reactions: [3]

(i) $CH_3CH_2Br \xrightarrow{\text{Alc.KOH}}$

(ii) $CH_3CH{=}CH_2 + O_3 \xrightarrow{\text{Zn/}H_2O}$

(iii) $CH_2{=}CH_2 + H_2O + [O] \xrightarrow{\text{Dil.}KMnO_4}$

9. (i) What are the conditions for a compound/species to be aromatic according to Huckel's rule? [3]

(ii) How will you convert Benzene to acetophenone?

10. (i) An alkene 'A' of molecular mass '28u' on treatment with bromine gives a product 'B'. The Compound 'B' on further dehalogenation with zinc gives back 'A'. Give the structures of 'A' and 'B' and also the sequence of reactions. [5]

(ii) Why is benzene extraordinarily stable although it contains three double bonds?

(iii) How can we convert ethyne into benzene?

UNIT TEST-II

Time Allowed: 1 Hr. **Maximum Marks : 20**

General Instructions:

(i) All questions are compulsory.

(ii) Maximum marks carried by each question are indicated against it.

---

1. The methyl group in Benzene ling is [1]
   - (a) Ortho directing (b) Ortho and meta directing
   - (c) Para directing (d) Ortho and para directing
2. Action of acetylene on dilute H_2SO_4 gives : [1]
   - (a) Acetic acid (b) Acetone
   - (c) Acetaldehyde (d) Ethyl alcohol
3. What is the other name for Geometrical isomerism? [1]

In the following questions a statement of Assertion (A) followed by Reason (R) is given. Use the following key to select correct answer :

(a) Both Assertion and Reason are correct but Reason is the correct explanation of Assertion.

(b) Both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

(c) Both Assertion and Reason are incorrect.

(d) Assertion is not correct but Reason is correct.

4. **Assertion:** Benzene does not decolourise bromine water.

   Reason: Benzene is stabilised by resonance due to delocalisation of π electrons.

5. **Assertion:** Acidity of C-H bond decreases in the order:

   $HC \equiv CH > H_2C = CH_2 > H_3C - CH_3$

   Reason: Greater the percentage s-character, more is the acidity of C—H bond.

6. How will you convert methyl bromide to ethane? [2]
7. Explain with the help of equation ozonolysis of propene. [2]
8. Write the mechanism of nitration of benzene. [3]
9. How will you convert benzene into [3]
   - (i) Acetophenone
   - (ii) Toluene
10. Write short note on the following : [5]
    - (i) Wurtz reaction.
    - (ii) Kolbe's electrolysis
    - (iii) Ozonolysis

*****

QUESTION FOR PRACTICES-I

1. What are the oxygen moles in 0.5 mol of $CaCO_3$?

   (a) 1 mol (b) 0.2 mol

   (c) 1.5 mol (d) 3.0 mol

2. What is the unit of wave number (ν) ?

3. The general configuration of 'f' block is

   (a) (n-1) f^{1-14} nd^{0-1} ns2 (b) (n-1)f^{0-1} nd^2 ns^2

   (c) (n-2) f^{1-14} (n-1) d^{0-1} ns^2 (d) (n-2) f^{1-14} (n-1)d^{0-2} ns^{0-1}

4. The shape of IBr^-_2 is

   (a) Tetrahedral (b) Planar

   (c) Linear (d) V-shape

5. The body temperature of a normal healthy person is 37ºC. Calculate its value in ºF.

6. Cs shows maximum photoelectric effect. Why?

7. Which of the following orbitals are possible?

   1s , 2s, 2p and 3f

8. $\Delta_f H^0$ for Graphite is ____________ .

Q.9 - Q.10 Assertion-Reason Type Questions

Each question contains statements-1 (Assertion) and Statement-2 (Reason) Examine the statements carefully and mark the correct answer according to the instruction given below:

A. If both the statements are true and statement-2 is the correct explaination of statement-1.

B. If both the statements are true but statement-2 is not the correct explanation of statement-1.

C. If statement-1 is true and statement-2 is false.

D. If statement-1 is false and statement-2 is true.

9. Statement-1 For reaction $A + B \rightleftharpoons C$, K = 4 on addition of catalyst K becomes more than 4.

   Statement-2 Catalyst only helps to attain the equilibrium faster from either end of reaction.

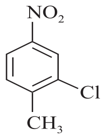
10. Statement-1 $Cl_2 + 2OH^- \longrightarrow ClO^- + Cl^-$ is a disproportionation reaction.
Statement-2 In disproportionation, the same element get oxidised as well as reduce.

11. –COOH is ortho, Para directing or Meta directing group? Give reason.

12. Draw Cis and trans isomers of CHCl=CHCl

13. At what temperature the entropy of a perfectly Crystalline solid is zero?

14. Write the IUPAC name of following



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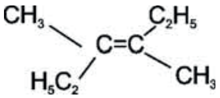


15. In what period and group will an element with Z = 118 will be present.

16. Calculate number of atoms in 52u of He.

17. Which series of lines of the hydrogen spectrum lie in the visible region?

18. Write the name of element with highest electron gain enthalpy.

19. Draw the shape of ClF_3.
OR, Draw the shape of SF_6.

20. What is bond order of O_2 molecule

21. (i) What are the number of waves made by a Bohr electron in an orbit of maximum magnetic quantum number 3?
(ii) If kinetic energy of a particle is doubled. What will happen to de Broglie wavelength as compared to previous de Broglie wavelength.

22. List the possible values for all the quantum numbers for the following sub shell
(i) 2p
(ii) 4f

23. (i) Second I.E. (Ionisation Enthalpy) is always more than first Ionisation energy.
(ii) Why first electron gain enthalpy of sulphur is more negative than oxygen.

24. Balance the following reaction by (ion-electron or oxidation number method)

$$Cl_2O_7(g) + H_2O_2\ (aq.) \longrightarrow ClO_2^-(aq.) + O_2\ (g) + H^+\ \text{(Acidic medium)}$$

25. Calculate oxidation state of Fe in Fe_3O_4.

26. A sample of 0.5g of an organic compound was treated according to Kjeldahl's method. The ammonia evolved was absorbed in 50 ml of 0.5 M H_2SO_4. The residual acid required 60 mL of 0.5 solution of NaOH for neutralisation. Find the percentage composition of nitrogen in the compound.

27. (i) Out of benzene, m-dinitrobenzene and toluene, which will undergo nitration most easily and why?
    (ii) What effect does branching of an alkane chain has on its boiling point?

28. Give IUPAC name the following:
    (i) $C_6H_5\ CH_2CH_2OH$
    (ii) $(CH_3)_2\ CH_2CH_2CHO$

29. Inductive effect is of permanent nature while electromeric effect is only temporary. Explain.

30. A crystalline salt on being rendered anhydrous loses 45.6% of its weight. The percentage composition of the anhydrous salt is
    Al = 10.50%, K= 15.1%, S= 24.96%, O=49.92%
    Find the simplest formula of the anhydrous and crystalline salt.
    (Atomic Mass : K= 39, Al = 27, S = 32, O = 16).

31. Explain following with example :
    (i) Aufbau's Rule (ii) Hund's Rule (iii) Pauli's Exclusion Principle

32. An ion with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

33. Explain the bonding in SF_6 using hybridisation concept and define what is hybridisation.

34. On the basis of molecular orbital theory find the bond order, molecular orbital configuration and magnetic nature of O_2^+.

35. Calculate pH of a solution obtained by mixing 50 mL of 0.2 MHCl with 49.9 mL of 0.2 m NaOH solution.

36. Explain Born Haber cycle with by considering example of formation of $MgCl_2$ as given in the chemical reaction

$$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$$

37. Calculate the standard enthalphy of formation of $CH_3OH(l)$ from the following data:

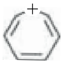
$$CH_3OH(l) + \frac{3}{2}O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta_C H^o = -726 \text{ kJ mol}^{-1}$$

$$C(g) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta_f H^o = -393 \text{ kJ mol}^{-1}$$

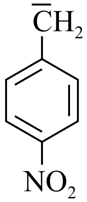
$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta_f H^o = -286 \text{ kJ mol}^{-1}$$

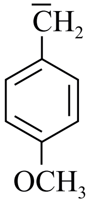
38. Arrange following in order of increasing stability

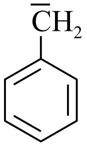
(a) $\overset{+}{C}H_3$ $\qquad$ $CH_3\overset{+}{C}H_2$ $\qquad$ $(CH_3)_3C^+$

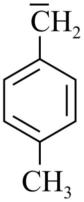


(b)









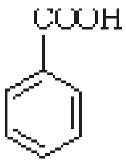
(c) $CH_3\dot{C}H_2, \quad \dot{C}H_3, \quad CH_2-CH=\dot{C}H_2, \quad \dot{C}H=CH_2$

39. (i) What is ambident nucleophile? Mention one example.

(ii) Distinguish between homolytic and hetrolytic bond cleavage.

(iii) Which one is stronger acid and why

H—COOH,



40. (a) What is a buffer solution? Give example.

(b) What is common ion effect?

(c) Define Le-Chatlier principle and explain effect of following :

(i) Change of concentration (ii) Change of pressure

41. (a) Find out K_c for following reaction

$2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$; $K_p = 1.8 \times 10^{-4}$ at 500K

(b) $K_p = 0.04$ atm at 899K. What is the equilibrium concentration of C_2H_6 where it is placed in a flask at 4.0 atm pressure and allow to come to equilibrium

$C_2H_6 \rightleftharpoons C_2H_4(g) + H_2(g)$

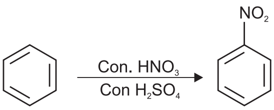
(c) What is the unit of K_p for the following chemical reaction?

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

42. A compound contains 4.07% H, 24.27%C and 71.65% Cl. Its molar mass is 98.96 g. What are its empirical and molecular formula.

43 Calculate the uncertainty in the position of a dust particle with mass equal to 1 mg if the uncertainity in its velocity is 5.5×10^{-20} ms^{-1}.

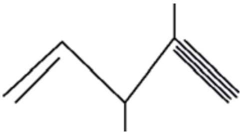
44. (i) Explain with the help of mechanism



(ii) Draw the Newman structure of (a) 2-Methyl butane

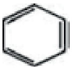
(b) 1-Methyl prop-1-ene

(iii) Calculate total number of σ and π bond(s) in

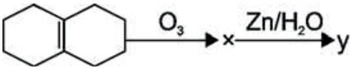


45. (i) Explain Kolbe's electrolysis with mechanism.

(ii) State Huckel Rule's.

Check whether  is an aromatic or non aromatic, anti-aromatic.

(iii) Write the product



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QUESTION FOR PRACTICES-II

1. The number of nodal planes in p_x orbitals is

(a) 1 (b) 2 (c) 3 (d) 0

2. Which of the following has smallest bond angle?

(a) H_2O (b) H_2S (c) NH_3 (d) SO_2

3. For a reaction to be spontaneous at all the temperature:

(a) ΔG –ve, ΔH +ve, ΔS +ve (b) ΔG +ve, ΔH -ve, ΔS +ve

(c) ΔG –ve, ΔH -ve, ΔS - ve (d) ΔG –ve, ΔH -ve, ΔS +ve

4. The enthalpy of all elements in their standard states are:

(a) unity (b) zero

(c) less than zero (d) different for each element

5. Which of the following has largest ionic radii?

(a) Na^+ (b) Mg^{2+} (c) F^- (d) O^{2-}

6. When carbon is bonded to four other atoms or groups it uses _________ hybrid orbitals.

7. Surface tension _____________ with increase in temperature.

8. The second electron gain enthalpy is _________.

Directions for Question No. **9 and 10 :** A statement of assertion (A) followed by a statement of reason (R) is given. Choose the correct option out of the choices given below for each question:

(a) A and R both are correct and R is the correct explanation of A.

(b) A and R both are correct but R is not the correct explanation of A.

(c) A is true but R is false.

(d) A and R both are false.

9. Assertion: The entropy of ice is less than that of water.

Reason: Ice has a cage like structure.

10 Assertion: London forces are much more stronger between Xenon atoms than between Helium atoms.

Reason: Xenon atom is bigger than Helium atom.

11. Write empirical formula of CH_3COOH and K_2CO_3

12. Define mole fraction.

13. Mention the quantum number which determines the energy of electron in the H-atom.

14. How many unpaired electrons are there in Ni^{2+} ion? (Given : Z =28)

15. State the condition for the formation of precipitate.

16. Write the conjugate acid and conjugate base of HSO_4^-.

17. Using VSEPR theory draw the shape of XeF_4 molecule.

18. Write IUPAC name of $CH_2 = CH — CH(OH)C \equiv CH$.

19. In a reaction between an oxidant and a reductant which will give up electrons and which will accept electrons?

20. Calculate oxidation number of Cr in $K_2Cr_2O_7$ and Mn in $KMnO_4$.

21. Write the conjugate acids for the following Bransted bases:

$NH_2^-, NH_3, HCOO^-$

22. What will be the mass of one ^{12}C atom in grane?

23. How are frequency and wave number related to each other.

24. Why the I.E. of Ga is greater than Al?

25. What effect does branching of an alkane chain has on its boiling point?

26. Cis But-2-ene has lower melting point that trans But-2-ene. Give reason.

27. Balance the following redox reaction in acidic medium by ion electron method.

$$Zn(aq.) + NO_3^- (aq.) \longrightarrow Zn^{2+} (aq.) + N_2O (g) + H_2O$$

28. (i) Mg^{2+} ion is smaller than O2-ion, although both are isoelectronic. Give reason.

(ii) Write IUPAC name and symbol for the element with atomic no. 120.

29. (i) Mention the number of radial nodes in 6s orbitals.

(ii) Write electronic configuration of Fe^{2+} ion. (Given, Z = 26)

30. Calculate the wavelength of a ball of mass 0.1kg moving with a velocity of $10ms^{-1}$. (Given, $h = 6.626 \times 10^{-34}$ Js)

31. Arrange Benzene, *n*-Hexane and Ethyne in decreasing order of acidic behaviour.

32. Which of the two $O_2NCH_2CH_2O^-$ or $CH_3CH_2O^-$ is expected to be more stable and why?

33. Write the difference between precision and accuracy.

34. Out of molarity and molarity which changes with temperature.

35. (i) Stability of carbocations follows the order 3° > 2° > °1. Explain this order of stability of carbocations.

(ii) In what manner is Electromeric effect different from Inductive effect?

36. (i) Mention the reason of not using Wurtz reaction for the preparation of unsymmetrical alkanes from alkyl halides.

(ii) How will you convert Benzene to p-Nitrobromobenzene?

37. The density of 3M solution of NaCl is 1.25g/mL. Calculate the molality of the solution. (Given: Atomic masses: Na=23u , Cl=35.5u)

38. Calculate the molarity of nitric acid (HNO_3) in a sample having a density 1.41g/mL and mass percent of nitric acid in it being 69%. (Atomic mass: N=14u, H=1u, O=16u)

39. (i) The ball hit with a hockey by a player does not form a wave. State reason.

(ii) Write the possible values of 'm' for an electron with l=2.

(iii) Chromium has configuration $3d^5 4s^1$ and not $3d^4 4s^2$.Explain.

40. (i) Explain non linear shape H_2S and non-planar shape of PCl_3 using VSEPR theory.

(ii) Can we have a diatomic molecule with its ground state molecular orbitals full with electrons. Give reason for your answer.

41. Calculate enthalpy change for the reaction:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

The enthalpy of formation of $CH_4(g)$,$CO_2(g)$ and $H_2O(l)$ are $-74.8 kJmol^{-1}$, $-393.5\ kJmol^{-1}$ and $285.8 kJmol^{-1}$ respectively.

42. Calculate the bond enthalpy of Cl-Cl bond from the following data:

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl\ (l) + HCl \quad \Delta H = -100.3 kJmol^{-1}$

Given: bond enthalpies of C — H, C — Cl and H — Cl bonds are 413, 326 and 431 $kJmol^{-1}$ respectively.

43. A neon-dioxygen mixture contains 70.6 g dioxygen and 167.5 g neon. If pressure of the mixture of gases in the cylinder is 25 bar, what is the partial pressure of dioxygen and neon in the mixture? (Atomic mass: O = 16u, Ne = 20u)

44. An organic compound contains 69% carbon and 4.8% Hydrogen, the remainder being oxygen. Calculate the masses of carbon dioxide and water produced when 0.20 g of this substance is subjected to complete combustion.

45. (i) State the necessary compound to be aromatic according to Huckel's rule.

(iii) Explain why alkyl groups act as electron donors when attached to a π system.

46. (i) Draw the resonance structures of Phenol.

(ii) Suggest a method used to purify the liquids which have high boiling points and decompose below their boiling points.

47. (i) Explain common ion effect with example.

(ii) The concentration of hydrogen ion in a sample of soft drink is 4 x 10-3 M. Calculate its pH.

(iii) What is the effect of removal of CH_3OH on the equilibrium of the reaction, $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$.

48. (i) Calculate H_3O^+ ion concentration of a water sample having pH = 6.78.

(ii) Define buffer solutions.

(iii) State Lewis definition of acids and bases. Give one example of each.

49. (i) An alkene A on ozonolysis gives a mixture of propanal and pentan-3-one. Write the structural formula of A.

(ii) Complete the following reactions:

(a) $CH_3CH{=}CH_2 + HBr \xrightarrow{\text{organic peroxide}}$

(b) $CH_3CH_2Br + Na \xrightarrow{\text{Dry Ether}}$

(iii) Write a short note on Friedel Crafts alkylation.

50. (i) Why eclipsed form and staggered form of ethane cannot be isolated at room temperature?

(ii) State Markovnikov's rule.

(iii) Which out of Ethene or Ethyne is more acidic and why?

(iv) What happens with 2-Bromobutane is being treated with KOH (alcoholic)?

****

QUESTION FOR PRACTICES-III

1. Which of the following d-orbital has double disc (body shooter shape)

(a) d_{xy} (b) d_{z^2} (c) $d_{x^2-y^2}$ (d) d_{yz}

2. Among the following, the compound that contains, ionic, covalent and Co-ordinate linkage is

(a) NH_4Cl (b) NH_3 (c) NaCl (d) CO_2

3. Which of the following is a state function

(a) q (b) w (c) q + w (d) None of these

4. Which one of the following thermodynamic quantities is not a state function

(a) Gibb free energy (b) Enthalpy

(c) Entropy (d) Work

5. The correct order of electron gain enthalpy with negative sign of F, Cl, Br and I having atomic number 9, 17, 35 and 53 respectively is

(a) I > Br > Ci > F (b) F > Cl > Br > I

(c) Cl > F > Br > I (d) Br > I > Cl > F

6. Arrange the following carbonations in order of their increasing stability

$(CH_3)_3C^+$, $CH_3CH_2^+$, $(CH_3)_2CH^+$, CH_3^+

7. The IUPAC name and symbol of element with Z = 120 is __________.

8. Direction for Question No. 9 to 10 : A statement of assertion (A) followed by a statement of Reason (R) is given.

(i) A and R both are correct, and R is correct explanation of A.

(ii) A and R both are correct, but R is not correct explanation of A.

(iii) A is true but R is false.

(iv) A and R both are false.

9. Assertion (A) : Enthalpy of graphite is lower than that of diamond.

Reason (R) : Enthalpy of graphite is greater than that of diamond.

10. Assertion (A) : Lower the critical temperature of the gas; more easily can it be liqufied.

Reason (R) : Critical temperature is the temperature above which a gas cannot be liquefied depending upon the pressure.

11. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how much volume of water vapour would be produced?

12. How is pH scale affected by increasing temperature.

13. Why are alkali metals used in photoelectric cells?

14. Write electronic configuration of Cr^{3+} ion. (Atomic No. of Cr = 24)

15. Draw the shape of ClF_3 molecule according to VSEPR theory?

16. Write the conjugate acid and conjugate base of H_2O.

17. Write the relation between solubility and solubility product of $Cr_2(SO_4)_3$.

18. Consider the given standard reduction potentials of following elements A, B, C & D and arrange them in order of their increasing reducing power.

$A = -3.71V, \quad B = -0.76V, \quad C = +0.34\ V, \quad D = +0.80\ V$

19. For the reaction $Co(g) + Cl_2(g) \rightleftharpoons CoCl_2(g)$ the value of kc/kp is equal to.

20. Chemical species having identical bond order have same bond dissociation enthalpies.

21. Write IUPAC name of $CH_3 — CO — CH_2 — CHO$.

22. Out of Cis-But-2-ene and Trans-but-2-ene which has greater boiling point and why?

23. Neither '*q*' nor '*w*' is state functions but $q + w$ is a state function. Explain.

24. (i) What is the lowest value of n that allows 'g' orbitals to exist?

(ii) Why 4s orbital is filled before 3d?

25. Calculate wave number for the longest wavelength transition in the Balmer Series of hydrogen atom.

26. Explain why?

(i) $\Delta_i H_1$ of 'N' is more than that of 'O'.

(ii) A cation is smaller than parent atom.

27. Balance the following redox reaction in acidic medium by ion electron method.

$$Cr_2O_7^{2-}(aq.) + SO_2(g) \longrightarrow Cr^{3+}(aq.) + SO_4^{2-}(aq.)$$

28. What is chromatography? Name different types chromatographic processes.

29. Write reaction is suitable for the preparation of both symmetrical and unsymmetrical alkanes.

30. In sulphur estimate, 0.157 g of an organic compound gave 0.4813 g of barium sulphate. What is the percentage of sulphur in the compound?

31. Write chemical equations to covert:

(i) Ethyne to Ethanal

(ii) Benzene to m-Nitrotoluene

32. The density of 3 M solution of NaCl is 1.25 g ml^{-1} calculate molarity of the solution.

33. 3 g of H_2 react with 29 g of O_2 yield H_2O

(a) Which is the limiting reagent?

(b) Calculate the maximum amount of H_2O that can be formed.

34. Calcium carbonate reacts with aqueous HCl according to the reaction:

$$CaCO_3(s) + 2HCl(aq.) \longrightarrow CaCl_2(aq.) + CO_2(g) + H_2O(l)$$

What mass of CaCO3 is required to react completely with 25 mL of 0.75 M HCl?

35. (i) State Hund's rule of maximum multiplicity of electrons.

(ii) The mass of an electron is 9.1×10^{-31} kg. If its K.E. 3.0×10^{-25} J, calculate its wavelength.

36. N_2 is diamagnetic while O_2 is paramagnetic. Explain on the basis of molecular orbital theory.

37. Explain the structure of PCl_5 according to hybridization. Why all P — Cl bonds.

38. What will be the pressure exerted by a mixture of 3.2 g of methane and

4.4 g of carbon dioxide contained in 9 dm^2 flask at 27°C.

39. Calculate the compressibility factor for CO_2, if one mole of it occupies 0.4 litre at 300K and 40 atm. Comment on the result.

40. The combustion of 1 mol of benzene takes place at 298K. After combustion CO_2 and H_2O are formed and 3267 kJ mol^{-1}, $\Delta_f H^{\ominus}(H_2O) = -393$ kJ mol^{-1}.

41. For the reaction : $2A(g) + B(g) \longrightarrow 2D(g)$, $\Delta U^{\ominus} = -10.5$ kJ and $\Delta S° = -44.1$ J K^{-1}. Calculate $\Delta G°$ for the reaction, and predict whether the reaction will occur spontaneously.

42. Account for the following:

(i) Halogen have very high negative electron gain enthalpy

(ii) Ionisation enthalpy of nitrogen (z = 7) is more than oxygen (z = 8)

43. (i) Out of $NO_2 — CH_2 — O$ and $CH_3 — CH_2 — O^-$ which is more stable and why?

(ii) Why is it necessary to prepare Lassaigne extract for detection of N, S and halogens?

(iii) Define the term hyperconjugation.

44. (i) State Le Chatelier's principle.

(ii) Calculate the pH of 10^{-8} M HCl.

(iii) A sample of HI (g) is placed in flask at a pressure of 0.2 atm. At equalibrium the partial pressure of HI (g) is 0.04 atm. What is the Kp for given equilibrium?

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

45. (i) Define common ion effect.

(ii) Write the relationship between K_p and K_c for the reaction:

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

(iii) Equal volume of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate? (For cupric iodate $K_{sp} = 7.4 \times 10^{-8}$)

46. Compare relative stability of following species and predict their magnetic properties:

$O_2, O_2^+, O_2^-, O_2^{2-}$

47. Give reason for the following:

(i) NH_3 has higher boiling point than pH_3

(ii) Ionic compound do not conduct electricity in solid state

(iii) H_2O has bent structure.

48. (i) Propanal and pentan-3-one are the ozonolysis product of an alkene. What is the structural formula and IUPAC name of alkene?

(ii) Give the main products of the reactions:

(a) $C_6H_6 + CH_3Cl \xrightarrow{\textbf{Anhydrous } \mathbf{AlCl_3}}$

(b) $CH_3Cl + Na \xrightarrow{\textbf{Dry ether}}$

(c) $CH_3 — CHCl — CH_2 — CH_3 + KOH(alc.) \longrightarrow$

49. (i) Give chemical euqation for each of the following :

(a) Decarboxylation

(b) Friedel Craft acetylation

(ii) Addition of HBr to propene yields 2-Bromopropane, while in the presence of benzoyl peroxide, the same reaction yields 1-Bromopropane. Explain and give mechanism.

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QUESTION FOR PRACTICES-IV

1. Number of matamers possible for molecular formula C_4H_{10} are ________.

2. How is the mass of an element related the number of atoms present in it.

3. Haemoglobincontain 0.25% iron by mass. The molecular mass of haemoglobinis 89600. Calculate the number of iron atoms per molecule of haemoglobin.

4. Orbital which is not possible

(a) 2p (b) 3d

(c) 3s (d) 3f

5. Which of the following is paramagnetic

(a) CO (b) O^-_2

(c) N_2 (d) NO^+

6. Which of the following is not correct?

(a) ΔG is zero for a reversible reaction.

(b) ΔG is positive for a spontaneous reaction.

(c) ΔG is negative for a spontaneous reaction.

(d) ΔG is positive for a non-spontaneous reaction.

Direction for Question No. 7 and 8 :

(i) If both assertion & reason are true and the reason is the correct explanation of the assertion.

(ii) If both assertion & reason are true but the reason is the correct explanation of the assertion.

(iii) If assertion is true statement but reason is false.

(iv) If both assertion and reason are false statements.

7. Assertion : Number of orbitals in 3rd shell is 9.

Reason : Number of orbitals for a particular value of n = n^2.

8. Assertion : Ionic radius of Na^+ is smaller than Na.

Reason : Effective nuclear charge of Na^+ is higher than Na.

9. What is the IUPAC name of picric acid.

10. Why noble gases have positive electron gain enthalpy?

11. How many molecules of SO_2 are present in 11.2 L at STP?

12. Calculate the number of atoms in 35 g of Li (Atomic mass of Li = 7 u).

13. For an isolated system, $\Delta U = 0$, what will be ΔS ?

14. Write the statement of third law of thermodynamics.

15. Calculate the number of sigma and pie bonds in C_2H_2?

16. Give the relation between K_a and K_b.

17. Write the conjugate base of H_2CO_3.

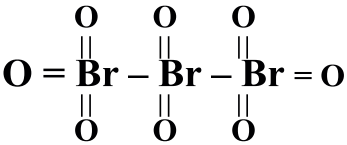
18. Write the oxidation number of Cr in $K_2Cr_2O_7$.

19. Write the stock notation MnO_2.

20. Suggest angular and spherical node on (a) 4s (b) 2p (c) 3d

21. Name the series of the line spectrum of H-spectrum which lies in visible region.

22. Assign the oxidation number of underline element



$$O = \underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Br}} - \underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Br}} - \underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{Br}} = O$$

23. Identify oxidizing and reducing agent in:

$$CH_4 + 4Cl_2 \longrightarrow CCl_4 + 4HCl.$$

24. Out of stagged and eclipsed form which is were stable and why?

25. What are carbocations? Give the stability order of different carbocations.

26. Explain why :

(a) The three electron present in 2p subshell of nitrogen remain upaired.

(b) Cr has configuration $3d^5\,4s^1$ and not $3d^4\,4s^2$.

27. Calculate the radius of Bohr's fifth orbit for hydrogen atom.

28. (a) Why electron gain enthalpy of Cl is more –ve than F.

(b) Give group number, period and block of an element having atomic number :

(i) 86 (ii) 55 (iii) 74 (iv) 63

29. Balance the following redox reaction in acidic medium

$$MnO_4^- + H^+ + Fe^{2+} \longrightarrow Mn^{2+} + H_2O + Fe^{3+}.$$

30. Write the resonance structure of (a) CO_3^{2-} (b) NO_3^-

31. Explain why $BeCl_2$ molecule have zero dipole moment.

32. 0.40 g of an organic compound gave 0.3 g of silver bromide of Carius method. Find the percentage of bromine in the compound. [Atomic mass of Ag & Br are 105 and 80 u]

33. Propanal and pentan-3-one are the ozonolysis product of an alkene. What is the structural formula of alkene?

34. Convert the following :

(a) 1-Bromopropane to 2-Bromopropane

(b) Ethanol to Glycol

35. Explain the following terms (a) Tautomerism (b) metamerism.

36. $CH_3CH{=}CH_2 + O_3 \longrightarrow X \xrightarrow{Zn/H_2O} A + B.$

Identify A and B. Also give reactions.

(a) Green house effect (b) Green chemistry

37. In a compound $C_xH_yO_z$, the mass % of C and H is 6 : 1 and the amount of oxygen present is equal to the half of the oxygen required to react completely C_xH_y. Find the empirical formula of the compound.

38. When light with a wavelength of 400 nm falls on the surface of sodium, electrons with a kinetic energy of 1.05×10^5 J mol^{-1} are emitted.

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

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

39. (a) Describe the hybridisation in case of C_2H_2.

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

40. (a) Use molecular orbital theory to predict why the He_2 molecule does not exist?

(b) Compare the stability of O_2 and O_2^+ on the basis of molecular theory.

41. Calculate the maximum work of expansion of two moles of an ideal gas when it expand isothermally and reversibly from 20 L to 40 L at 27ºC (R = 1.987 cal per kelvin per mol.

42. Heat produced during the combustion of benzoic acid (C6H5COOH) at 298 K and 1 atm is –2500 KJ per mole. What is ΔE for the reaction.

43. The standard enthalpy of combustion of Surcose $(C_{12}H_{32}O_{11})$ at 298 K producing $CO_2(g)$ and $H_2O(l)$ is $-5200.7 \text{ kJ mol}^{-1}$. If $\Delta_f H°$ of $CO_2(g)$ and $H_2O(l)$ are $-393.0 \text{ kJ mol}^{-1}$ and -286 kJ mol^{-1}.

44. Explain the First law of thermodynamics and how can you prove the following relation :

$\Delta H = \Delta U + P\,\Delta V$ from the first law of the thermodynamics $(\Delta U = q + w)$?

45. Give structural formula of following compound:

(a) But-2-enoic acid

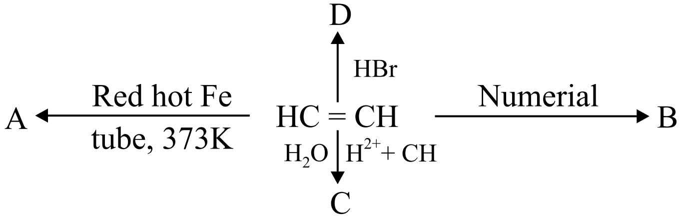
(b) 4-Hydroxypent-3-enoic acid

(c) 3-Cyclopropyl-3-methyulbutanal.

46. Explain the resonance effect and draw the resonance structures of $CH_2 = CH - CHO$. Indicate the relative stability of the contributing structures.

47. (a) Addition of HBr to propene yields-2-bromopropane, while in the presence of benzoylperoxide, the same reaction yields 1-bromopropane. Explain and give tis mechanism.

(b) Identify A, B, C & D



(c) Give a chemical test between ethen and ethane.

48. (a) Out of benzene and toluene, which will undergo nitration more easily and why?

(b) Explain the following :

(i) Friedal - Craft reaction

(ii) Wurtz reaction

49. (a) What is solubility product? How is it different from ionic product?

(b) Equal volume of 0.02 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate? (For cupric iodate Ksp = 7.4×10^{-8}).

50. (a) Define Le-Chatelier principle.

(b) Explain ionic product of water. What is the effect of temperature on ionic product of water?

(c) Calculate the pH of 10^{-10} M NaOH solution.

51. If 20 mL of 0.1 M $BaCl_2$ is mixed with 15 mL of 0.01 M $AgNO_3$, Predict the precipitation of AgCl if its solubility product is 1.47×10^{-10}.

52. Describe the effect of

(a) addition of Cl_2

(b) addition of NO_2Cl

(c) removal of NO_2

(d) removal of NO_2Cl

On the equilibrium of reaction:

$Cl_2(g) + 2NO_2(g) \rightleftharpoons 2NO_2Cl(g)$

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PRACTICE PAPER– I
CLASS - XI
CHEMISTRY THEORY (043)

Time: 3 Hours Max. Marks : 70

General instructions: read the following instructions carefully.

1. There are 33 questions in this question paper with internal choice.
2. Section A consists of 16 multiple choice questions carrying 1 mark each.
3. Section B consists of 5 short answer questions carrying 2 marks each
4. Section C consists of 7 short answer questions carrying 3 marks each.
5. Section D consists of 2 case-based questions carrying 4 marks each.
6. Section E consists of 3 long answer questions carrying 5 marks each.
7. All questions are compulsory.
8. Use of log table and calculator is not allowed.

SECTION-A

The following questions are multiple choice questions with one correct answer. Each question carries 1 mark. There is no internal choice in this section.

1. Among the following group, the group having highest priority in IUPAC nomenclature scheme is

(a) F (b) $-OCH_3$

(c) $-CH_3$ (d) $-COOH$

2. Which of the following is heavier in weight

(a) 2g He (b) 22.4L at STP He

(c) 20 moles of H_2 (d) 10 moles of N_2

3. The molarity of a solution obtained by mixing 750 ml of 0.5M HCI with 250ml of 2M HCI will be

(a) 0.975M (b) 0.875M

(c) 1.00M (d) 1.175M

4. What transition in H-spectrum would have the same wavelength as the Balmer transition n=4 to n=2 in the He^+ spectrum

(a) n=4 to n= 1 (b) n=3 to n=2

(c) n=3 to n=1 (d) n=2 to n=1

5. The total number of orbitals associated with third shell will be

(a) 2 (b) 4 (c) 9 (d) 3

6. The correct order of ionic radii of the species N^{3-}, O^{2-}, Na^+ and F^- is

(a) $Na^+ < F^- < O^{2-} < N^{3-}$

(b) $F < O^{2-} < N^{3-} > Na^+$

(c) $O^{2-} < N^{3-} < F^- > Na^+$

(d) $N^{3-} < Na^+ < F^- > O^{2-}$

7. Which molecule have zero dipole moment

(a) Ammonia

(b) water

(c) Carbon dioxide

(d) $CHCI_3$

8. Which of the following molécule has coordinate bond

(a) CH_4 (b) H_2O (c) CO_2 (d) NH_4^+

9. The correct order of the bond order in the following species are

(a) $O_2^{2+} > O_2^+ > O_2^-$

(b) $O_2^{2+} < O_2^- < O_2^+$

(c) $O_2^+ > O_2^- > O_2^{2-}$

(d) $O_2^- < O_2^+ < O_2^{2+}$

10. A process will be spontaneous at all temperature if

(a) $\Delta H > 0$ and $\Delta S < 0$

(b) $\Delta H < 0$ and $\Delta S > 0$

(c) $\Delta H < 0$ and $\Delta S < 0$

(d) $\Delta H > 0$ and $\Delta S > 0$

11. According to second law of thermodynamics

(a) $\Delta S_{total} = +ve$

(b) $\Delta S_{total} = -ve$

(c) $\Delta S_{system} = +ve$

(d) $\Delta S_{system} = -ve$

12. In what manner increase in pressure affect the following equilibrium?

$C(s) + H_2O(l) \rightleftharpoons CO(g) + Hg$

(a) Shift in forward direction

(b) Shift in reverse direction

(c) Increase in yield of hydrogen

(d) No effect

In the following questions (Q. 13 to 16) two statements are given labelled as Assertion (A) and Reason (R), choose the correct answer out of the following choices.

(a) Both A and R are true and R is correct explanation of A.

(b) Both A and R are true and R is not the correct explanation of A.

(c) A is true but R is false

(d) A is false but R is true

13. Assertion: ammonia is a lewis base

Reason: NH_3 molecule is pyramidal in shape.

14. Assertion: a liquid crystallises in solid and is accompanied by decrease in entropy

Reason: In crystals molecules organise in an ordered manner

15. Assertion: Oxidation state of hydrogen in HCI is +1 but in NaH is –1

Reason: NaH is a metal hydride and for hydrides hydrogen is assigned an oxidation state of –1.

16. Assertion: Benzene is non-aromatic compound

Reason: compounds which follow the Huckle's rule are aromatic in nature

SECTION B

17. The uncertainty in momentum of a particle is 2.5×10^{-14} gcms^{-1} with what accuracy can its position be determined. ($h = 6.25 \times 10^{-27}$ gcms^{-1})

18. Account for the following

(i) Nitrogen has more first ionization energy than oxygen.

(ii) The electron gain enthalpy of Cl is more negative than that of fluorine.

19. Calculate the pH of 0.02M acetic acid solution. K_a for acetic acid solution is 1.8×10^{-5} at 298K

20. (A) Give IUPAC name of $CH_2{=}CH\text{-}CH_2\text{–}CHOH\text{-}CH_3$

(B) give metamer of N-Ethylethanamine ($C_2H_5\text{–}NH\text{–}C_2H_5$)

OR

Explain (a) $ClCH_2\text{–}COOH$ is more acidic than $CH_3\text{–}COOH$

(b) arrange in order of increasing stability

$(CH_3)_3\,C^+$, $CH_3CH_2^+$, $CH_3C^+HCH_3$, CH_3^+

21. Explain the following (a) kolbe's electrolysis

(b) wurtz reaction

SECTION-C

22. Answer the following

(a) Why cation is always smaller than parent atom

(b) Why lonisation energy of B is less then Be

(c) Give group number and period mumber of an element having atomic number 75

23. (a) molar heats of combustion of $C_2H_2(g)$, C(graphite) and $H_2(g)$ are 310.62 kcal, 94.05 kcal and 68.12 kcal respectively. Calculate the standard heat of formation of $C_2H_2(g)$.

(b) Define intensive property

24. (a) for a reaction $2A + B \longrightarrow C$

$\Delta H = 400$ KJ mol^{-1} and $\Delta S = 0.2$ kJ K^{-1} mol^{-1}

At what temperature will the reaction become spontaneous considering ΔH and ΔS to be combatant over the temperature range.

(c) Define enthalpy of formation

25. At a certain temperature equilibrium constant (K_c) is 16 for the reaction

$SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$

If the container contain 1M concentration of each component initially then what is concentration of SO, at equilibrium.

26. Balance the following equation by in oxidation number method

$P_4(s) + OH^-(aq) \longrightarrow PH_3 + H_2PO_2^-$ (Basic medium)

27. (a) identify the electrophiles among the following

$:CCl_2$, NH_3, Br^+, OH^-

(b) Give any two difference between inductive and electromeric effect.

28. One mole of a hydrocarbon A reacts with 1 mole of bromine giving a dibromo compound B, $C_5H_{10}Br_2$. Compound A on treatment with cold dilute $KMnO_4$ solution form a compound C, $C_5H_{12}O$. On ozonolysis A gives equimolar quantities of propanone and ethanal. Deduce the structure of A and give reaction of A with $KMnO_4$ and ozonolysis.

SECTION D

29. Head the case study given below and answer the following questions:

The valence Shell Electron Pair Repulsion theory by Sidgwick and powell in 1940, proposed a simple theory based on the repulsive interaction of the electron pairs in the valence shell of the atoms. It was further developed and redefined by NYholm and Gillespie (1957) and they put forward the concept of important difference of lone pair and the bonding pairs of electrons. While the lone pairs are localized on the central atom, each bonded pair is shared between two atoms. In the molecules the bond pair, Lone pairs (if any) will occupy such position around the central atom to obtained a minimum repulsion energy states. These repulsion effects result in deviation from idealized shapes and alteration in bond angles in molecules.

(A) Which have higher bond angle NH_3 or H_2O. Give reason

OR

Give shape of following compounds on Basis of VSPER theory

(a) PCl_5 and (b) XeF_2

(B) Give an example of molecule having tetrahedral geometry

(C) Arrange in order of increasing repulsion

Lone pair-bond pair, bond pair-bond pair, lone pair-lone pair

30. Alkenes undergoes addition reaction and oxidation reactions due to presence of unsaturation Addition reactions are those in which the attacking reagent adds to the substrate molecule without elimination. ie. two molecules combine to form a single molecule. Such reactions are given by those molecules which possess at least one multiple bond. In such reactions for each π bond of the molecules, two sigma bond is formed and hybridization state of the C-atom changes from sp^2 to sp^3. This addition may be electrophilic addition, nucleophilic addition or free radical addition. Addition takes place either according to markownkov's rule or according to antimarkownikov's rule depending on molecule to be added or conditions of additions

(A) Complete the following reaction

$CH_3–CH=CH_2+HBr \longrightarrow$

$CH_3–CH=CH_2+HBr \xrightarrow{\text{peroxide}}$

An organic compound X of molecular formula C_6H_{12} absorb one mole of bromine to give $C_6H_{12}Br_2$. On ozonolysis it gives two molecules of propanal: Give structure of X and its reaction with ozone.

(B) How will you convert 2-bromopropane to 1-bromopropane.

(C) Name the monomer of polyethene.

SECTION-E

31. (a) calculate the ratio of energy of electron in the ground state of hydrogen to electron in the first excited state of Be^{+3}.

(b) Give electronic configuration of i) Cr ii) Fe^{+2}

(c) an electron is moving with velocity 2×10^3 mS^{-1} find the wavelength of the electron (h=6.62×10^{-34} js)

32. Give answer to the following questions:

(a) State markownikov's rule. Explain with example.

(b) Draw sawhorse projection of conformation of ethane.

(c) Halogens are ortho and para directing but deactivating. Explain.

OR

(a) What is peroxide effect. Explain with example.

(b) State homolytic fission.

(c) In sulphur estimation, 0.157 g of an organic compound gave 0.4813g of $BaSO_4$. What is percentage of sulphur in organic compound.

33. (A) a compound containing sodium, sulphur, hydrogen and oxygen gave the following results on analysis; Na 14.28% S=9.92% H=6.2%. Calculate the molecular formula of the anhydrous compound. If H is present in the form of water of crystallization only, what is the structure of the crystalline salt? (molecular weight of crystalline salt=322)

(B) why do we prefer molality over molarity for expressing the concentration of a solution?

(C) State limiting reagent

OR

(a) Calculate mole fraction of benzene in a solution containing 30% by mass in CCl_4.

(b) What volume of oxygen at NTP is needed to cause the complete combustion of 200ml of acetylene? Also calculate the volume of CO_2 formed,

Solution

1. (d) **2.** (d) **3.** (b) **4.** (d) **5.** (c)

6. (b) **7.** (c) **8.** (d) **9.** (a) **10.** (b)

11. (a) **12.** (b) **13.** (b) **14.** (a) **15.** (a)

16. (d)

17. $\Delta x = h/4 \times 3.14 \Delta p = 2.11 \times 10^{-14}$

18. (i) due to stable exactly half filled p^3 configuration in nitrogen

(ii) because of interelectronic repulsions in F due to small sire some of energy is used to overcome the repulsion

19. $[H^+] = \sqrt{K_a.C} = \sqrt{1.8 \times 10^{-5} \times 0.02} = 6 \times 10^{-4}$

$pH = -\log [H^+] = -\log (6 \times 10^{-4}) = 3.22$

20. (a) Pent-4-en-2-ol. (b) $CH_3–NH–C_3H_7$

OR

(a) $ClCH_2COOH$ is more acidic than $CH_3–COOH$ due to –1 effect of Cl

(b) $CH_3^+ < CH_3CH_2^+ < CH_3C^+HCH_3 < (CH_3)_3\ C^+$,

21. (a) $CH_3COONa \rightarrow CH_3COO^- + Na^+$

At anode: $CH_3COO^- \rightarrow CH_3CO\dot{O} + e^-$

$CH_3CO\dot{O} \rightarrow \dot{C}H_3 + CO_2$

$\dot{C}H_3 + \dot{C}H_3 \rightarrow CH_3 – CH_3$

At cathode: $2H_2O + e^- \rightarrow H_2 + 2OH^-$

(b) $CH_3Cl + 2Na + CH_3Cl \rightarrow CH_3\text{-}CH_3 + 2NaCl$

22. (a) due to increase of effective nuclear charge with lose of electron.

(b) due to more penetration of s electron to be removed from Be than B in which p electron is removed.

23. (a) $\Delta H = 2 \times (–94.05) + (–68.32) – (–310.62) = 54.20$ kcal

(b) the properties which depend on nature of substance and not on quantity of substance.

24. $\Delta G = \Delta H – T\Delta S$ for spontaneous process $\Delta G < 0$

i.e. $\Delta H – T\Delta S < 0 \quad \Rightarrow \quad \Delta H < T\Delta S \quad \Rightarrow \quad T > \Delta H/\Delta S$

$\Delta T > 400/0.2 \quad \Rightarrow \quad > 2000K$

25.

	$SO_2(g)$	+	$NO_2(g)$	$\rightleftharpoons$	$SO_3(g)$	+	$NO(g)$
T = 0	1		1		1		1
$T = t_{eq}$	1–X		1–x		1–x		1–x

Therefore $k = 16 = [1+x]^2/[1-x]^2$ Therefore x=0.6

Hence $[SO_2] = 1 – 0.6 = 0.4$ moles/Litre

26. $P_4(s) + 3OH^-(aq) + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$

27. (a) $:CC_2$, Br^+

(b) Inductive effect 1. It is permanent displacement of electron of C-chain due to presence of an atom or group of atom attached with C chain having electronegativity different from C.

2. Presence of multiple bond is not required

Electromeric effect: 1 it is temporary shifting of electrons of rt bond towards one of the bonded atom

2. Presence of multiple bond is required. Or any two.

28. A = $(CH_3)_2\ C{=}CHCH_3$

$(CH_3)_2C{=}CHCH_3 + H_2O + [O] \xrightarrow{KMnO_4/OH–} (CH_3)_2C(OH)\text{-}CH(OH)CH_3$

$(CH_3)_2\ C{=}CHCH_3, \xrightarrow{(i)\ O_{3}\text{–}\ (ii)\ Zn/water} (CH_3)_2C{=}O + O{=}CHCH_3$

29. (A) NH_3

Reason: NH_3 Have only one lone pair while water have two lone pair of electron. Due to higher lone pair-lone pair repulsion in water bond angle decrease.

OR

(a) PCl_5 trigonal bipyramidal

(b) XeF_2 linear

(B) Methane

(C) bond pair-bond pair < Lone pair-bond pair < lone pair-lone pair

30. (a) $CH_3–CH{=}CH_2 + HBr \longrightarrow CH_3–CHBr–CH_3$

$CH_3\text{-}CH{=}CH_2 + HBr \xrightarrow{peroxide} CH_3\text{-}CH_2\text{-}CH_2Br$

$X = CH_3–CH_2–CH{=}CH–CH_2–CH_3$

$CH_3–CH_2–CH{=}CH–CH_2CH_3 \xrightarrow{(i)\ O_{3}\text{–}\ (ii)\ Zn/water} CH_3–CH_2–CHO + OCH–CH_2–CH_3$

(b) $CH_3–CHBr–CH_3 \xrightarrow{KOH/alcohol} CH_3–CH{=}CH_2 \xrightarrow{HBr/peroxide} CH_3CH_2CH_2Br$

(c) ethene

31. (a) $E_1/E_2 = Z_1^2n_2^2/Z_2^2n_1^2 = 1/4$

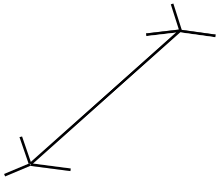
(b) $Cr = [Ar]3d^54s^1\ Fe^{+2} = [Ar]3d^6$

(c) $\lambda = h/mv = 6.62 \times 10^{-34}/9.1 \times 10^{-31} \times 2 \times 10^3 = 0.364 \times 10^{-6}m = 364nm$

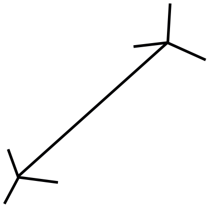
32. (a) when an unsymmetrical molecule is added to unsymmetrical alkene Then negative part of molecule goes to the C-atom which has less number of H-atom

e.g. $CH_3–CH{=}CH_2 + HBr \longrightarrow CH_3–CHBr–CH_3$

(b)



Eclipse



Staggered

(c) Halogens are ortho directing due to +R effect but deactivating due to effect.

OR

(a) When HBr is added to an unsymmetrical alkene in presence of peroxide then negative part (Br) goes to the C-atom which has more number of H-atom.

eg $CH_3–CH=CH_2 + HBr \xrightarrow{\text{peroxide}} CH_3–CH_2–CH_2Br$

(b) when a covalent bond is cleaved in such a way that each atom retain one electron of the shared pair, the fission of called homolytic fission.

(c) % of Sulphur = 32 × weight of $BaSO_4$ × 100/233 × weight of organic compound

= 32 × 0.4813 × 100/233 × 0.157

= 42.1%

33. (A)

Element	%age	atomic mass	relative no. of moles	whole no.ratio of atoms
Na	14.28	23	14.28/23=0.62	0.62/0.31=2
S	9.92	32	9.92/23=0.31	0.31/0.31=1
H	6.2	1	6.2/1=6.2	6.2/0.31=20
O	69.6	16	69.6/16=4.35	4.35/0.31=14

Empirical formula = $Na_2SH_{20}O_{14}$

Empirical formula mass 23 × 2 + 32 + 1 × 20 + 16 × 14 – 322

n = molecular formula mass/ empirical formula mass = 322/322 = 1

molecular formula = (Empirical formula)n = $(Na_2SH_{20}O_{14})1 = Na_2SH_{20}O_{14}$

since all hydrogen is present in form of water hence molecular formula is $Na_2SO_4 . 10H_2O$

(b) Because molality is independent of temperature while molarity decrease with increase of temperature.

(c) The reactant which is completely consumed in reaction is called limiting reagent.

OR

(a) Moles of benzene = 30/78=0.38 moles

Moles of CCl_4 = 70/154 = 0.45 moles

Total moles 0.38 + 0.45 = 0.83

Mole fraction of benzene = moles of benzene/ total moles = 0.38/0.83=0.45.

(b)	C_2H_2	+	$5/2O_2$	→	$2CO_2 + H_2O$	
	22400ml		56000ml		44800ml	at NTP
	200 ml		?		?	

Volume of oxygen required = 56000 × 200/22400 = 500ml

Volume of CO_2 produced = 44800 × 200/22400 = 400 ml

PRACTICE PAPER– II
CLASS - XI
CHEMISTRY THEORY (043)

Time: 3 Hours Max. Marks : 70

General Instructions: Read the following instructions creafully

1. There are 33 questions in the questions paper with internal choice.
2. Section A consists of 16 multiple-choice questions carrying 1 mark each.
3. Section B consists of 5 short answer questions carrying 2 marks each.
4. Section C consists of 7 short answer questions carrying 3 marks each
5. Section D consists of 2 case based questions carrying 4 marks each.
6. Section E consists of 3 long answer questions carrying 5 marks each.
7. All questions are compulsory
8. Use of log tables and calculators is not allowed.

SECTION-A

1. One mole of oxygenn gas at STP is equal to ....

(a) 6.022×10^{23} molecules of oxygen

(b) 6022×10^{23} atoms of oxygen

(c) 16 g of oxygen

(d) 32 g of oxygen

2. In which of the following pairs, the ions are iso electronic?

(a) Na^+, Mg^{2+} (b) $Al^{3+}, 0^-$

(c) Na^{2+}, O^{2-} (d) N^{3-}, Cl^-

3. The radius of which of the following orbit is same as that of first obit of hydrogen atom?

(a) He^+ (n = 2) (b) Li^{2+} (n = 2)

(c) Li^{2+} (n = 3) (d) Be^{3+} (n = 2)

4. Which of the following pairs of d-orbitals have electron density along the axis?

(a) dz^2, dxz (b) dxz, dyz

(c) dz^2, dx^2y^2 (d) dzy, $dx^2 - y^2$

5. The electronic configuration of gadolinium (Atomic No.64) is

(a) [Xe] $4f^3\,5d^5\,6s^2$ (b) [Xe] $4f^7\,5d^2\,6s^1$

(c) [Xe] $4f^7\,5d^1\,6s^2$ (d) [Xe] $4f^8\,5d^6\,6s^2$

6. The types of hybrid orbitals of nitrogen in NO_2^+, NO_3^- and NH_4^+ respecting are expected to be

(a) sp, sp^3 and sp^2 (b) sp, sp^2 and sp^3

(c) sp^2, sp and sp^3 (d) sp^2, sp^3 and sp

7. $\Delta_f U$ of formation of $CH_4(g)$ at certain temperature is –393 kJ mol^{-1}. The value of $\Delta_f H$ is

(a) zero (b) $< \Delta_f U$

(c) $> \Delta_f U$ (d) equal to $\Delta_f U$

8. 0.1M CH_3COOH is 1.34%, ionised, calculate its ka

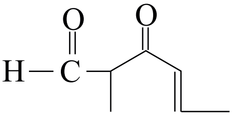
(a) 1.8×10^{-5} (b) 1.8×10^{-4}

(c) 5×10^{-4} (d) 4×10^{-5}

9. Which of the following compound contain all the carbon atoms Same hybridisation state?

(a) $H–C{\equiv}C–C{\equiv}C–H$ (b) $CH_3–C{\equiv}C–CH_3$

(c) $CH_2{=}C{=}CH_2$ (d) $CH_2{=}C{=}C{=}CH_2$

10. The IUPAC name of the compound 

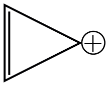
(a) 5-formyl hex-2-en-3-one

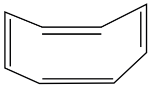
(b) 5-methyl-4- oxohex-2–en-5-al

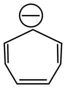
(c) 3-keto-2 methyl hex-5-enal

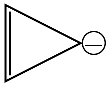
(d) 3-keto-2-methyl hex-4-enal

11. Which of the following structure is atomatic?

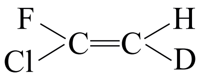
(a) 

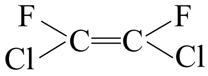
(b) 

(c) 

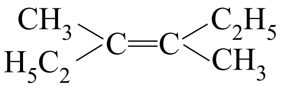
(d) 

12. Which of the following will not show geometrical isomenim 1

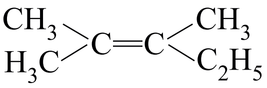
(a) 
$(F)(Cl)C{=}C(H)(D)$

(b) 
$(F)(Cl)C{=}C(F)(Cl)$

(c)



(b)



In the following questions (Q.13-Q.16), a statement of assertion followed by a statement of reason is given. Choose the Correct answer out of the following choices

(a) Both Assertion & Reason are true & Reason is correct explanation of the Assertion

(b) Both Assertion and Reason are true but Reason is not correct explanation of Assertion.

(c) Assertion is a true but Reason in false

(d) Assertion is false but Reason is true

13. Assertion. It is impossible to determine the exact position and exact momentum of an electron simultaneously.

Reason: The path of an electron in an atom is clearly defined.

14. Assertion: Boron has a smaller 1st Ionisation enthalpy than Beryllium

Reason: The penetration of a 2s electron to the nucleus is more than the 2p electron hence 2p electron is more shielded by the inner core of electron than the 2s electron.

15. Assertion: Heat capacity is the amount of heat required to raise temperature of body by 1K.

Reason: Heart capacity is an extensive property and it depends upon the size of the body.

16. Assertion: Black body is an ideal body that units and absorb all radiations.

Reason: The frequency of radiation emitted by a body goes from a lower frequency to higher frequency with an increase in temperature.

SECTION B

This section antains 5 questions with internal choice in one question. The following questions are very short answer type and carry 2 mark each.

17. (i) Write the significant figures in the following:

(a) 2.00 (b) 4.8046

(ii) Write the electronic configuration of Cu(29).

18. How will you convert following:

(i) Propene to 2, 3-Dimethyl butane

(ii) Ethyne to Ethanal

19. If 50g of $CaCO_3$ is treated with 50g of HCl, how many grams of CO_2 can be produced according to following equations:

$CacO_3(a) + 2HCl\ (dil) \longrightarrow CaCl_2\ (aq) + CO_2(g) + H_2O(l)$

20. The wavelength of 1st spectral line in the Balmer series is 6561 A°. Calculate the wavelength of the 2^{nd} spectral line in Balmer series.

Or

The radius 1st Bohr's orbit of hydrogen atom is 0.529 A°. Calculate the radius of—

(i) the 3^{rd} orbit of He^+ ion

(ii) the 2^{nd} orbit of Li^{2+} ion

21. What are values of 'n' & 'I' for valence election of Na(11).

SECTION C

This section contains 7 questions with internal choice in one question. The following questions are short answer type and carry 3 marks each.

22. (i) Define Markonikov's Rule. Illustrate with the help of an example.

(ii) Distinguish b/w 1-Butyne & 1-butene

(iii) Write all possible isomers of C_4H_9Br.

23. (i) State Heisenbergs uncertainty principle.

(ii) Calculate the uncertainity in the position of an election if uncertainty in its velocity is 0.001%.

$(m = 9.1 \times 10^{-31}\ kg,\ h = 6.63 \times 10^{-34} Js)$

24. (i) Identify the oxidising and reducing agent int he following reaction

$H^+ + H_2O \longrightarrow OH^- + H_2$

(ii) Balance the following equation-

$MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$

(iii) What is the oxidation state of Fe in $K_4\ [Fe(CN)_6]$?

25. (i) Define molar heat capacity

(ii) The enthalpy of combustion of benzoic acid at 298K and 1 atm is $-2546\ kJ\ mol^{-1}$. What is ΔU of the reaction)

26. (i) State Hund's Rule

(ii) Calculate the wavelength associated with a Proton moving with 1.0 × 10^3 ms^{-1}. (mass of proton = 1.67×10^{-27} kg, h = $6{\cdot}63 \times 10^{-34}$Js]

Or

(i) Why are Bohr's orbits called stationary state?

(ii) Which series of hydrogen spectrum lies in UV region?

(ii) Which of the following have highest frequency.

X-rays, γ-rays, Microware

27. A compound is made up of 2 elements A & B has A=70%., B=30%. Their relative number of moles in the compound is 1.25 & 1.88. Calculate:

(i) Atomic Masses of the elements A & B.

(ii) Molecular formula of the compound if molecular mass of compound is 160u.

28. (i) Define Le-chatelier's Principle

(ii) The Equilibrium constant for the reaction $H_2(g) + Br(g) \rightleftharpoons 2HBr$ at 1024 k is 1.6×10^5. Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into sealed container at 1024 k.

SECTION D

The following questions are case based questions. Each question has an internal choice and carries 4 marks each. Read the passage carefully and answer the following questions.

29. To locate the address of a person, we require name, house number, city State, pin code. Similarly the electron in an atom can be located with the help of quantum numbers. These quantum numbers describe the energy level of an orbital and define the shape and orientation of the region in space where the election will be found. The principal, azimuthal and magnetic quantum numbers come from the solution of schro-dinger wave equation. The fourth quantum number called spin quantum number represents the spin of the election about its own axis. Answer the following questions.

(a) How many electrons in the ground state of neon have $l + m_l = 0$

(Atomic No. of Ne = 10)

(b) Which quantum number distinguish between the 2 electrons present in the same obital

(c) Predict whether the following set of quantum numbers can exist or not for an electron.

(i) $n = 1, l = 0, m_l = 0, ms = -1/2$

(ii) $n = 5, l = 3, m_l = -4, ms = +1/2$

OR

What will be the maximum possible, no of electrons having $ms = -1/2$ for $n = 5$?

30. The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. The expermental approach required the introduction of new unit for amount of substances, the mole, which remains indispensable in modern chemical science. The mole is an amount unit similar to unit like pair, dozon. gross, etc. It provides a specific measure of the number of atoms of molecules in a bulk sample of matter. A mole is defined as the amount of substance containing the same number of discrete entities (atoms, molecules, ions etc.) as the number of atoms in a sample of pure 12(weighing exactly 12g one latin connotation for the word "mole" is large mass") or bulk" which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic propety, bulk mass and an extremely important fundamental property, number of atoms molecules and so forth. The number of entities composing a mole has been experimentally determined to be 6.022×10^{23}. 6.022×10^{23}, a fundamental constant named Avogaro's number (NA) or the Avogadro constant in honor of Italian scientist Amedco Avogactor. This constant is properly reported with an explicit unit of "per mole", a conveniently rounded version being 6.022×10^{23}/mol. Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different since the masses of the individual atoms are drastically different, the molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol).

(a) When an antacid tablet is used, $Ca(OH)_2$ reacts with HCl in the stomach to form inert $CaCl_2$ and H_2O. If the molar mass of $Ca(OH_2)$ is 75 g/mol, how many moles of HCl are required to fully react with 150g of $Ca(OH)_2$.

(b) Calculate no. of atoms of He present in 100u.

(c) How are 0.5 mol Na_2CO_3 and 0.5 M Na_2CO_3 different from each other?

OR

At STP, what will be the volume of 6.022×10^{23} molecules of H_2.

SECTION E

Following questions are long answer type and carry 5 marks each. All question have an internal choice.

31. (a) Equilibrium constant for a reaction is 100. What will be equilibrium constant for reverse reaction?

(b) Write the conjugate acids for bronsted base OH^- & CH_3COO^-

(c) Determine the pH of 10^{-8} M HCl solution taking into account the H+ produced by water also. (Log 11 = 1.04 M)

OR

(a) What is effect of temperature m value of Ksp?

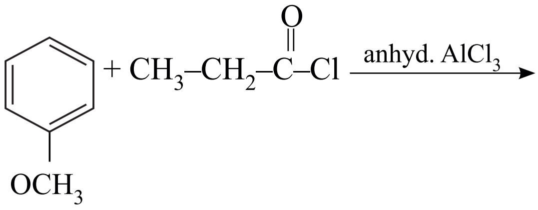
(b) Calculate pH of buffer solution Containing 0.01M solution of NH_4OH and 0.1 M solution of NH_4Cl. pKb of $NH_4OH = 4.75$.

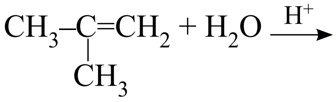
(c) Classify the following into Acidic and Basic Salt.

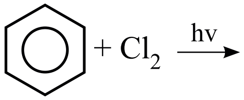
CH_3COONa, NH_4Cl, Na_2CO_3, $CuSO_4$

32. (a) Propanal and Propan-2-one are the ozonolysis products of an alkene what is the structural formula of the alkene.

(b) Give the major products of the reactions

(i)


(ii)

$$CH_3-\underset{\substack{|\\ CH_3}}{C}=CH_2 + H_2O \xrightarrow{H^+}$$

(iii)


OR

(a) An alkyl halide 'A' of the formula $C_6H_{13}Cl$ on treatment alc. KOH give two isomeric alkene (B) & (C) with formula C_6H_{12}. Both alkenes on hydrogenation give 2,3-Dimethylbutane Predict the structures of 'A', 'B' and 'C'.

(b) How is benzene obtained from ethyne! What is the name of the process?

(c) Convert 1-Propanol to 2-propanol.

33. (a) (i) Use Molecular orbital theory to predict why Be_2 molecule does nto exist.

(ii) Compare the stability of N_2^+ and N_2.

(b) (i) Why is bond length of all C=O bonds is equal in carbonate ion)

(ii) Why dipole moment of BF_3 is zero but NF_3 is not?

(iii) Why does water have boiling point higher than HF?

PRACTICE PAPER– III
CLASS - XI
CHEMISTRY THEORY (043)

Time: 3 Hours Max. Marks : 70

General Instructions: Read the following instructions carefully

1. There are 33 questions in the questions paper with internal choice.
2. Section A consists of 16 multiple-choice questions carrying 1 mark each.
3. Section B consists of 5 short answer questions carrying 2 marks each.
4. Section C consists of 7 short answer questions carrying 3 marks each
5. Section D consists of 2 case based questions carrying 4 marks each.
6. Section E consists of 3 long answer questions carrying 5 marks each.
7. All questions are compulsory

8 Use of log tables and calculators is not allowed

SECTION – A

1. Radius of the just Bohr's orbit of Hydrogen atom is ($a_0 = 52.9$ pm) :

(a) 26.3 pm (b) 13.15 pm

(c) 52.9 pm (d) 105.8 pm

2. The number of water molecules in a drop of water weighing 0.018 g is :

(a) 6.022×10^{26} (b) 6.022×10^{20}

(c) 6.022×10^{23} (d) 18.066×10^{23}

3. Splitting of spectial lines in an electric field is called.

(a) Zeeman effect (b) Shielding effect

(c) Photoelelctric effect (b) Strark effect.

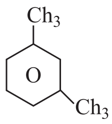
4. Which of the following has minimum band length.

(a) O_2 (b) 0_2^+ (c) O_2^- (d) O_2^{2-}

5. Which of the following is para magnetic.

(a) H_2 (b) H_2 (c) Li_2 (d) O_2

6. The IUPAC name of is.



(a) Dimethyl benzene (b) methyltoluene

(c) 1, 3–dimethyl benzene (d) 1,4 dimethylbenzene

7. When 22.4 L of H_2 is mixed with 11.2 L of cl_2 each at73 k at 1 atm the wales of Hcl formed used to :

(a) 2 moles of $Hcl_{(g)}$ (b) 1.5 moles of $Hl_{(g)}$

(c) 0.5 moles of $Hcl_{(g)}$ (d) 1 moles of $Hcl_{(g)}$

8. In which of the following pair both molecule do not possess same type of hybridisation.

(a) CH_4 and H_2O (b) Pd_5 and SF

(c) SFo and XeF_5 (d) BCl_3 and Ncl_3

9. The difference between H and U at constant volume is equal to :

(a) R (b) $P\,\Delta U$

(c) $V\,\Delta P$ (d) $\frac{3}{2}R$

10. The intermediate during the addition of HCl to propene in the propene of peroxide is:

(a) $CH_3^+CH_2Cl$ (b) $CH_3^+\,CHCH_3$

(c) $CH_3CH_2CH_2$ (d) $CH_3CH_2CH_2$

11. Oxygen has oxidation state of +2 in :

(a) H_2O_2 (b) H_2O

(c) OF_2 (d) CO_2

12. A process will be spontaneous all temperature is :

(a) $\Delta H > O$ and $S > O$ (b) $\Delta H > O$ and $S < O$

(c) $\Delta H < O$ and $S > O$ (d) $\Delta H < O$ and $S < O$

For question number 13 to 16 two statements are given are labelled as assertion (A) and other labelled as reason (R). Select the correct answer to these questions from the codes (a), (b), (c) and (d) as given below :

(a) Both Assertion (A) and reason are correct and Reason (R) is the correct explanation of the assertion (A)

(b) Both Assertion (A) and reason (R) are correct and reason (R) is not the correct explanation of the Assertion (A)

(c) Assertion (A) is correct but Reason (R) is incorrect.

(d) Assertion (A) is correct but Reason (R) is incorrect.

13. **Assertion (A) :** All isotopes of a given element show the same type of chemical properties.

Reason (R) : The chemical properties of an atom are controlled by the number of valence electron in the atom.

14. **Assertion (A) :** One mole of So_2 contains double the number of molecules present in one moles of 0_2.

Reason (R) : Molecular mass of So_2 is double to that of 0_2.

15. **Assertion (A) :** The atomic radii of Mg is smaller Na.

Reason (R) : Mg has a lower nuclear charge than Na.

16. **Assertion (A) :** A resonance hybrid is always more stable than any of its camonical structures.

Reason (R) : This stability is due to delocalization of electrons.

SECTION – B

17. Balance the following redox reaction :

$MnO4^- + I^- + H^+ \rightarrow Mn^{2+} I_2 + H_2O$

18. Give reason :

(i) The first lonisation enthalpy of nitrogen is greater than that of oxygen.

(ii) The electron gain enthalpies of Be and Mg are positive.

19. How can you convert.

(i) Ethane to ethanoic acid

(ii) iso-prepyl bravside to + bromo propane.

20. H° for the reaction

$2H_2(g) + O_2(g) \rightarrow 2H_2O\,(l)$ is -572 Kjmol^{-1}

Find $\Delta_j H^0$ for $H_2O\,(l)$

21. Draw Cis and trans isomers of hex-2-ene which isomer will have higher boiling point and why.

SECTION – C

22. Is the density of methanol (CH_3OH) is 0.793 kg L^{-1} What is its volume needed for making 2.5 L of its 0.25 M solution. (atomic mass : C = 124, H = 14, 0 = 164)

23. Describe the hybridiszation in pcl_5. Why are the axial bonds longer as compared to equitorial bonds.

24. Explaint the following :

(i) So_3 act as an electrophile.

(ii) Benzylic free radical is more stable than allylic free ladical.

(iii) $O_2NCH_2CH_2O^-$ is more stable than $CH_3CH_2O^-$

25. The solubility product of $Al(OH)_3$ is 2.7×10^{11} calculate its solubility in gL^{-1} and pH of this solution (atomic mass of Al = 274)

26. Complete the following reactions :

(i) $CH_2 = CH_2 + H_2O \xrightarrow[\Delta]{H^+}$

(ii) $CH_3 - CH = CH_2 + HBr \xrightarrow{\text{Peroxide}}$

(iii) $CH_3CH = CHCH_3 \xrightarrow[Zn/H_2O]{O_3}$

27. What volume of 0.6 m Hcl has enough Hcl to react exactly with 25ml of aq NaOH having concentration of 0.5 M?

28. (i) An atomic orbital has n = 3? What are the possible values of l and m?

(ii) List the quantam numbers (ml and l) of electron for 3d-orbit al

(iii) Which of the following orbitals are possible.

1p, 2s, 2p and 3s.

29. The phenomenon for ejection of electrons from the surface of metal when light of suitable frequency falls on it, is called the photoelectric effect and the emitted electrons are called photoelectrons.

The result observed in this experiment were

(i) The electrons are ejected from the metal surface as soon as the beam of light hits the surface.

(ii) The number of electrons ejected is proportional to the intensity or brightness of light.

(iii) For each metal, there is a characteristic minimum frequency, v_0 (called the threshold frequency) below which this effect does not observe. When frequency $v > v_0$, the ejected electrons come out with certain kinetic energy. The kinetic energies of these electrons increase as the frequency of the light used increases.

(a) The work function of a metal is 4.2 eu If radiation of 2000 A° fall on teh metal then the kinetic energy of the fastest photoelectron is ______

(b) What is the work function of the metal, n the light of wavelength 4000 A° generates photoelectron of velocity $6 \times 10^5 \text{ms}^{-1}$ from it?

(c) The number of electron ejected in the photoelectric experiment is proportional to the ______

Or

(c) Kinetic energy of the ejected electron is equal to ______

30. An oxidation-reduction (redox) is a type of chemical reaction that involves a transfer of electrons between two species. An oxidation-reduction is any chemical reaction in which the oxidation number of a molecule, atom, or ion changes by gaining or losing an electron. Redox reactions are common and vital to some of the basic functions of life, including photosynthesis, respiration, combustion, and corrosion or rusting. The oxidation state (OS) of an element corresponds to the number of electrons (e^-) that an atom loses, gains, or appears to use when joining with other atoms in compounds.

the oxidation state of an individual atom is 0. The total oxidation state of all atoms in : a neutral species is 0 and in an ion is equal to the ion charge. Group 1 metals have an oxidation state of +1 and group 2 an oxidation state of + 2. The oxidation state of fluorine is –1 in compounds. Hydrogen generally has an oxidation state of +1 in compounds. Oxygen generally has an oxidation state of –2 in compounds. In oxidation state of +1 in compounds. Oxygen generally has an oxidation state of –2 in compounds. In binary metal compounds, group 17 elements have an oxidation state of –1, group 16 elements of –2, and binary metal compounds, group 17 elements have an oxidation state of –1, group 16 elements of –2, and

(a) One mole of acidified $K_2Cl_2O_1$ an reaction with excess KI will liberate n mole of I_2 then the value of n is

(b) Determine the oxidation no of underlined atom in the following molecules.

(i) $Cr0_5$ (ii) $H_2P_40_7$

(c) When electrons are transferred from Zn to Cu^{2+} in copper sulphate the energy is

______

Or

(c) In alkaline medium Clo_2 oxidises H_2O_2 to O_2 and itself gets reduced to cl^-. How many moles of H_2o_2 are oxidised by 1 mole of clo_2?

31. (i) The longest wavelength doublet absorption transition is observe at 589 nm and 589.6 nm. Calculate the frequency of each transition and energy difference between two excited states.

(ii) In which atom, the outer most electron can have the following set of quantum numbers?

$n = 3, l = O, m_l = 0, m_s = -1/2$

Or

(i) The velocity associated with a proton moving in a potential difference of 1000 V is 4.37×10^5 ms^{-1}. If the hockey ball of mass 0.1 kg is moving with this velocity, calculate the wavelength associated with this velocity.

(ii) The radius of first Bohr orbit of hydrogen atom is 0.529 A. Claculate the radii of (a) the third orbit of He^+ ion and (b) the second orbit of Li^{2+} ion.

32. (i) The value of K_p for the reaction,

$CO_2(g) + C(s) \rightleftharpoons 2CO(g)$

is 3.0 at 1000 K. If initially

$P_{CO2} = 0.48$ bar and $p_{CO} = 0$ bar and pure graphite is present, calculate the equilibrium partial pressures of CO and CO_2.

(ii) $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g); DH = -117\ kJ$

(a) Predict the effect of an increase in concentration of NO on the equilibrium concentration of NO_2.

(b) Predict the effect of decrease in pressure as a result of increased volume on the equilibrium concentration of NO_2.

Or

(i) Calculate the pH of 0.08 M solution of hypochlorous acid, HOCl. The ionisation constant of the acid is 2.5×10^{-5}. Determine the percent dissociation of HOCl.

(ii) Ionisation constant of a weak base MOH, is given by the expression

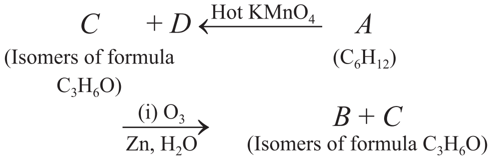
$$K_b = \frac{[M^+][OH^-]}{[MOH]}$$

Values of ionisation constant of some weak bases at a particular temperature are given below.

Base	**Dimethyl amine**	**Urea**	**Pyridine**
K_b	5.4×10^{-4}	1.3×10^{-14}	1.77×10^{-9}

Arrange the bases in decreasing order of the extent of their ionisation at equilibrium. Which of the above base is the strongest?

33. (i) Give the structures of A and B.



(ii) Arrange benzene, n-hexane and ethyne in decreasing order of acidic behaviour. Also give reason for this behaviour.

Or

An alkyl halide C_5H_{11} Br(A) reacts with alcoholic KOH to give an alkene 'B', which reacts with Br_2 to give a compound 'C', which on dehydrobromination gives an alkyne 'D'.

On treatment with sodium metal in liquid ammonia, one mole of 'D' gives one mole of the sodium salt of 'D' and half a mole of hydrogen gas. Calculate hydrogenation of 'D' yields a straight chain alkane. Identify A, B, C and D. Give the reactions involved.

COMMON ANNUAL SCHOOL EXAMINATION
(2024-25)
CLASS - XI
Sub. : CHEMISTRY (043)

Time: 3 Hours — Max. Marks : 70

General Instructions:

1. All questions are compulsory.
2. This question paper consists of 35 questions divided into five sections A, B, C, D and E.
3. Section A: Q1 to Q18 are objective type questions. Carrying 1 mark each.
4. Section B: Q19 to Q25 are very short answer type questions carrying 2 marks each.
5. Section C: Q26 to Q30 are short answer type questions carrying 3 marks each.
6. Section D: Q31 to Q32 are of passage/case study based questions carrying 4 marks each.
7. Section E: Q33 to Q35 are long answer type questions carrying 5 marks each.
8. There is no overall choice in question paper, only internal choices are available.
9. Use of calculator and log tables is not permitted.

SECTION – A

Question no. 1 to 16 are Multiple Choice Questions (MCQ), carrying one mark each. There is no internal choice in this section.

1. The total number of electrons in 18 mL of water (density = 1 g mL^{-1}) is :

(a) 6.02×10^{23} (b) 6.02×10^{24}

(c) 6.02×10^{25} (d) $6.02 \times 18 \times 10^{25}$

2. The wavelength of a ball of mass 0.1 kg moving with a velocity of 10 ms^{-1} is :

($h = 6.626 \times 10^{-34}$ Js)

(a) 6.626×10^{-33} m (b) 6.626×10^{-35}

(c) 6.626×10^{-34} m (d) 10 m

3. What are the maximum number of orbitals that can be identified with the following quantum number?

$n = 3, = 1$

(a) 4 (b) 3 (c) 2 (d) 1

4. Correct IUPAC symbol for the element with atomic number 118 is :

(a) Uue (b) Une (c) Uno (d) Uno

5. Which of the following molecule is a polar molecule?

(a) $BeCl_2$ (b) BF_3 (c) NH_3 (d) CH_4

6. In the structure of ClF_3, the number of lone pair(s) of electrons on central atom 'Cl' is:

(a) Four (b) Three (c) Two (d) One

7. Reaction, A+B C + D + q is found to have a positive entropy change. The reaction will be :

(a) possible at high temperature (b) possible only at low temperature

(c) not possible at any temperature (d) possible at any temperature

8. A thermodynamic system in which there is no exchange of matter but exchange of energy is possible, is called _____ system.

(a) Adiabatic (b) Isloated

(c) Open (d) Closed

9. What is the change in oxidation number of chlorine in the following reaction?

$$N_2H_{4(l)} + ClO^-_{3(aq)} \rightarrow NO_{(g)} + Cl^-_{(g)}$$

(a) +5 to –1 (b) +7 to –1

(c) +3 to –1 (d) +3 to + 1

10. Homolytic fission lead to the formation of _____.

(a) Nucleophile (b) Free radical

(c) Carbocation (d) Carbanion

11. Which of the following is an electrophile?

(a) BF_3 (b) H_2O

(c) CN^- (d) OH^-

12. Which of the following has zero dipole moment?

(a) But-1-ene (b) Cis-But-2-ene

(c) Trans-But-2-ene (d) 2-Methylprop-1-ene

Answer the questioins 13 to 16, two statement are given - one labelled as Assertin (A) and the other labelled as Reason (R) Select the correct answer to these questions from the codes (a), (b), (c) and (d) as given below :

(a) Both Assertion (A) and Reason(R) are true and Reason (R) is the correct explanation of the Assertion (A).

(b) Both Assertion (A) and Reason (R) are true, but Reason (R) is not the correct explanation of the Assertion (A).

(c) Assertion (A) is true, but Reason (R) is false.

(d) Assertion (A) is false, but Reason (R) is true.

3. **Assertion (A) :** lg O_2 and lg O_2 have an equal number of atoms.

Reason (R) : Mass of 1 mole atoms is equal to its gram atomic mass.

4. **Assertion (A) :** It is impossible to determine the exact position and exact momentum of an electron simultaneously.

Reason (R) : The path of an electron in an atom is clearly defined.

15. **Assertion (A) :** A liquid crystalises into a solid and is accompanied by drecrease in entropy.

Reason (R) : In crystals, molecules organise in an ordered manner.

16. **Assertion (A) :** A free radical is paramagnetic species.

Reason (R) : A free radical is formed in homolytic fission of covalent bond.

SECTION-N

The following consists of 5 short answer type questions. Each question carries 2 marks. There is internal choice in one question.

The empirical formula and molecular mass of a compound are CH_2O and 180 g mol^{-1} respectively. What will be the molecular formula of the compound?

Among the elements B, Al, C and Si, identify the element :

(a) Which was highest first ionisation enthalpy?

(b) Which was most negative electron gain enthalpy?

Calculate the number of kJ of heat necessary to raise the temperature of 60.0 g of aluminium from 35°C to 55°C. Molar heat capacity of Al is 24 J mol^{-1} k^{-1}. (Atomic mass Al = 27u)

OR

What will be the enthalpy change for the reaction?

$$H_{2(g)} + Br_{2(g)}\ 2HBr_{(g)}$$

Give that Bond energy of H_2, Br_2 and HBr is 435 kJ mol^{-1}, 192 kJ mol^{-1} and 368 kJ mol^{-1} respectively.

20. For the following reaction :

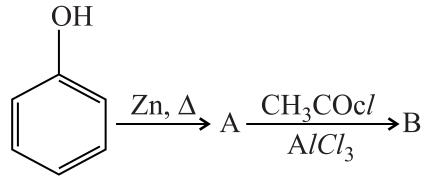
$$2A_{2(g)} + B_{(g)} \rightleftharpoons C_{(g)},\ \Delta H = -ve$$

What would be the effect on equilibrium if :

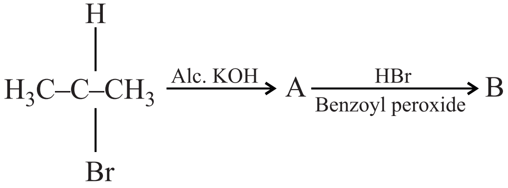
(i) more A_2 is added (ii) temperature is increased.

Write structure of A and B in the following reactions :

(a)



(b)



$$H_3C-\overset{\displaystyle H}{\underset{\displaystyle Br}{C}}-CH_3 \xrightarrow{\text{Alc. KOH}} A \xrightarrow[\text{Benzoyl peroxide}]{\text{HBr}} B$$

SECTION-C

The following consists of 7 short answer questions. Each question carries 3 marks. There is internal choice in one question.

22. 10.0 kg of $H_{2(g)}$ and 20.0 kg $Cl_{2(g)}$ are mixed to produce $HCl_{(g)}$. Calculate the amount of $HCl_{(g)}$ formed. Identify the limiting reagent in the production of $HCl_{(g)}$ in this situation.

23. (a) State Heisenberg Uncertainty Principle.

(b) If the position of the electron is measured within an accuracy of $\pm$ 0.002 nm. Calculate the uncertainty in the momentum of the electron. ($h = 6.626 \times 10^{-34}$ Js) 3

24. (a) List the quantum numbers (n and l) for the valence electron of Na atom (z = 11).

(b) Write a cation and an anion which is isoelectronic with Ar (z = 18)?

(c) Find the number of radial nodes for 3p orbital.

25. Account for the following :

(a) First ionization enthalpy of nitrogen is more than of oxygen.

(b) Noble gases have biggest atomic size in their respective period.

(c) Ionic radius of Na^+ is larger than Mg^{2+}.

26. (a) Give an example of disproportionation reaction.

(b) Balance the following redox reaction in acidic medium :

$$Cr_2O^{2-}_{7(aq)} + SO^{2-}_{3(aq)} \rightarrow Cr^{3+}_{(aq)} + SO^{2-}_{4(aq)}$$

27. (a) On complete combustion, 0.246 g of an organic compound gave 0.198 g of carbon dioxide and 0.1014 g of water. Determine the percentage of carbon in the compound. (Atomic mass : C = 12u, O = 16u)

(b) CCl_4 does not give precipitate of AgCl on heating with silver nitrate. Why?

28. (a) An alkene 'A' on ozonolysis gives a mixture of ethanal andpenta-3-one. Write structure and IUPAC name of 'A'.

(b) Give a chemical test to distinguish between but-1-yne and but-2-yne.

OR

(a) Which of the following compound(s) will show cis-trans isomerism?

(i) $CH_3\,CH{=}CH_2$ (ii) $CH_3CH = CH\,CH_3$

(b) How can you convert the following :

(i) Acetylene to nitrobenzene

(ii) Benzene to p-nitrotoluene

SECTION–D

The following questions are case based questions. Read the case carefully and answer the questions that follow :

29. Chemical or physical processes lead to evolution or absorption of heat (q), part of which may be converted into work (w). These quantities are related through the first law of thermodynamics. First law of thermodynamics does not guide us about the direction of chemical reactions i.e. what is the driving force of a chemical reaction? Another state function, S, entropy is a measure of disorder or randomness. For a spontaneous change, total entropy change is positive. Chemical reactions are generally carried at constant pressure, so we define another state function Gibbs free energy, G, which is related to entropy and enthalpy changes of system by the Gibbs equation.

Answer the following questions :

(a) On mole of acetone requires less heat to vapourise than I mol of water. Which of the two liquids has higher enthalpy of vapourisation?

(b) Heat and work done individually are not state functions but their sum is always a state function. Explain why?

(c) For the reaction at 298 K,

$2A + B \rightarrow C$

$\Delta H = 400$ kJ mol^{-1} and $S = 0.2$ kJ mol^{-1}. At what temperature will the reaction become spontaneous considering H and S to be constant over the temperature range?

OR

The reaction of cyanamide, $NH_2\,CN_{(s)}$, with dioxygen was carried out in a bomb calorimeter, and ΔU was found to be -742.7 kJ mol^{-1} at 298 K. Calculate enthalpy change for the reaction at 298K.

$$NH_2CN_{(s)} + \frac{3}{2}O_{2(g)} \rightarrow N_{2(g)} + CO_{2(g)} + H_2O_{(l)}$$

30. Petroleum and natural gas are the main sources of alkanes. Alkanes are almost non-polar molecules. They posses weak vander forces. Grease (mixture of higher alkanes) is non-polar and hence, hydrophobic is nature. Boiling points (b.p.) of some alkanes are given below :

Molecular formula	**Name**	**b.p. / K**
C_2H_6	Ethane	184.4
C_3H_8	Propane	230.9
C_4H_{10}	Butane	272.4
C_5H_{12}	Pentane	309.1
C_5H_{12}	2-Methylbutane	300.9
C_5H_{12}	2,2-Dimethylpropane	282.5

The important reactions of alkanes are free radical substitution, combustion, oxidation and aromatization. Alkanes show conformational isomerism due to free rotation about the C–C sigma bonds. Out of staggered and the eclipsed conformations of ethane, staggered conformation is more stable.

Answer the following questions :

(a) Draw the Newman's projections for staggered and eclipsed confrontations of ethane.

(b) Which one of the two conformations (staggered or eclipsed) of ethane is more stable and why?

(c) Give reasons :

(i) n-Butane has higher boiling point than ethane.

(ii) n-Pentane has highest boiling point among n-pentane, 2-Methylbutane and 2,2-Dimethylpropane.

Or

Complete the following reactions :

(i) $CH_3CH_2CH_2CH_2CH_2CH_3 \xrightarrow[\substack{773K \\ 10-20\,Atm}]{Cr_2O_3}$

(ii) $CH_3Br \xrightarrow{Na/Ether}$

SECTION – E

This section consists of 3 long answer type questions. Each question carries 5 marks. All questions have internal choice.

31. (a) Draw separate diagrams showing the formation of bonds and bonds in ethyne.

(b) Give reason for the following :

(i) O–O bond length in ozone molecule is 128 pm which is in between O-O single bond length 148 pm and O=) double bond length 121 pm.

(ii) Hydrogen fluoride exist as dimer.

(iii) Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that in ammonia.

OR

(a) Draw the shapes of the following molecules on the basis of VSEPR theory.

(i) NH_4^+ (ii) SF_4

(b) Write the molecular orbital configuation of O_2 molecule. Calcualte its bond order and precit its magnetic behaviour.

32. (a) State Le-Chatelier's principle.

(b) Define solubility product. Write solubility product expression in terms of molar solubility for $FeCl_3$.

(c) At 450 K, $K_p = 2.0 \times 10^{10}$ bar^{-1} for the given reaction at equilibrium.

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

What is K_c at this equilibrium? ($R = 0.0831$ L bar K^{-1} mol^{-1})

OR

(a) Explain common ion effect with one example.

(b) At 700 K, equilibrium constant for the reaction :

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$ is 50. What will be the equilibrium constant for the reverse reaction?

(c) Calculate the pH of 10^{-8} M HCl solution. ($\log 10 = 1, \log 11 = 1.02$)

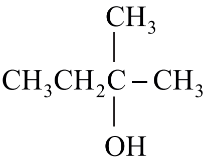
(d) Which of the following are Lewis acids?

H_2O, BF_3, H^+ and NH_3

33. (a) Explain the following terms :

(i) Inductive effect (ii) Carbocation

(b) Write down IUPAC name of the following compound :



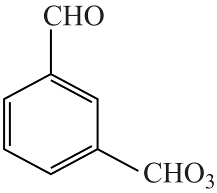
$$CH_3CH_2\overset{\displaystyle CH_3}{\underset{\displaystyle OH}{C}}-CH_3$$

Also write the structure of one of its position isomer.

(c) What is the state of hybridisation of underlined carbon atom in $H_2C = \underline{C} = CH_2$?

Or

(a) Identify the functional groups in the following compound :



(b) Draw the structure of Pent-4-en-2-ol

(c) Give reason :

(i) Alkyl groups act as electron donar when attached to a system.

(ii) $ClCH_2CH_2O^-$ is more stable than $CH_3CH_2O^-$

(d) Write the principle of chromatography.

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